Periodic Classification of Elements and Chemical Bonding

CHEMISTRY
GRADE 9

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Module - II

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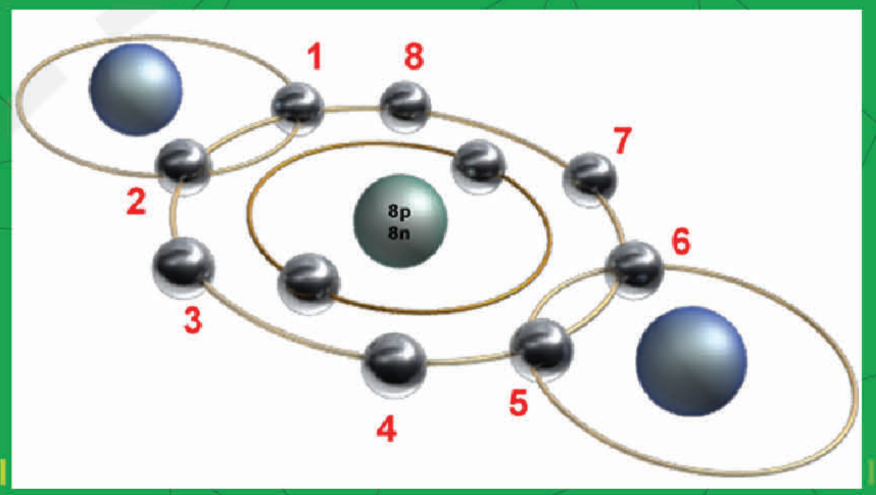

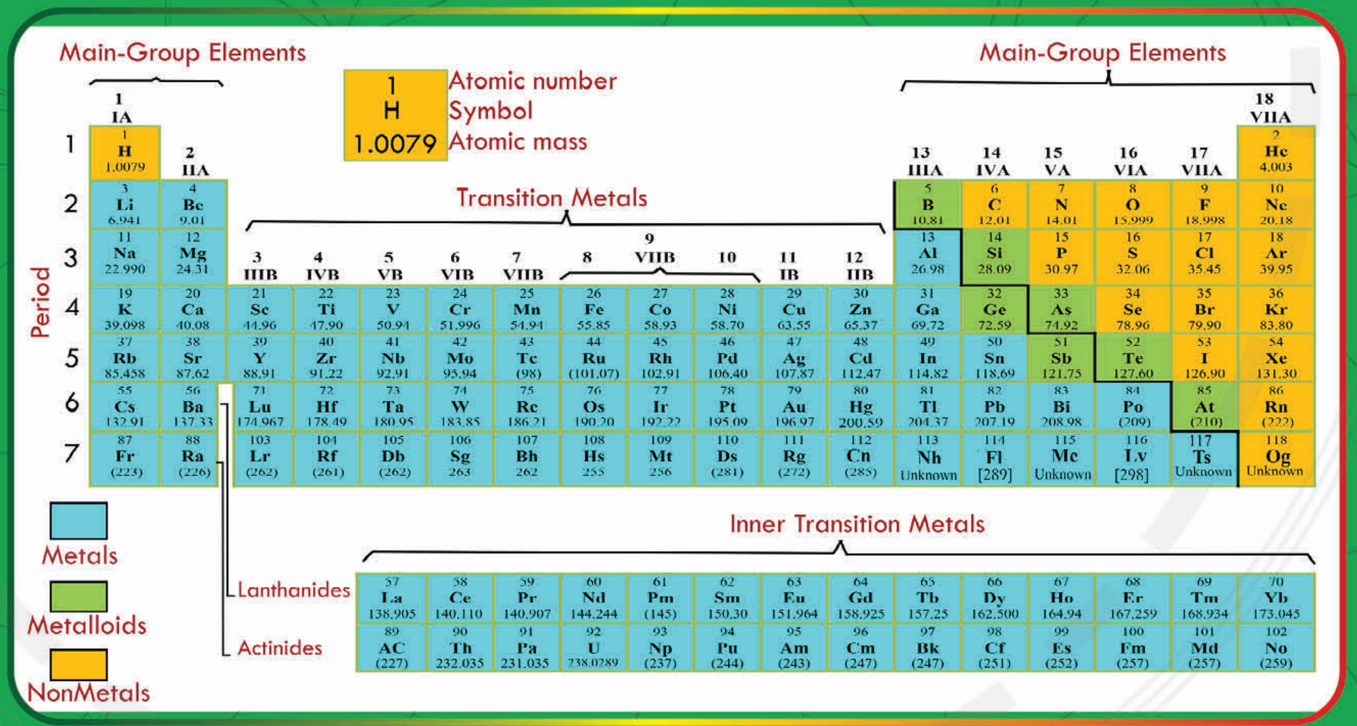

Federal Democratic Republic of Ethiopia
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MODULE - II
GRADE 9

Writers:
Sisay Tadesse (Prof.)
Tegene Tesfaye (Ph.D.)

Editors:
Anteneh Wasyhun (Ph.D.) (Curriculum Editor)
Kenenisa Beresa (M.A.) (Language Editor)

Illustrator:
Abinet Tilahun (M.Sc.)

Designer:
Konno B. Hirbaye (M.Sc.)

Evaluators:
Nega Gichile (B.Sc., M.A.)
Sefiw Melesse (M.Sc.)
Tolessa Mergo (B.Sc., M.Sc.)

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MODULE CONTENTS

There are a number of icons, or symbols in this teaching material. The meanings of each icon is as follows.

☸ This tells you there is an overview of the unit and what the unit is about.

? This tells you there is an in-text question to answer or think about in the text.

☞ This tells you to take note of or to remember an important point.

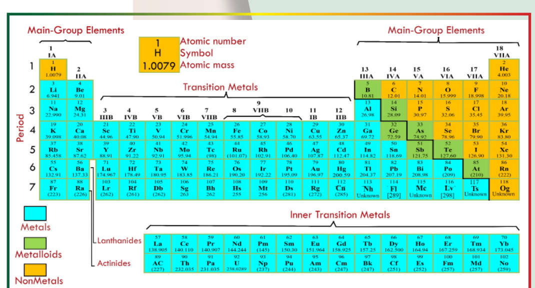
This tells you there is a self-test for you to do.

☑ This tells you there is a checklist.

✍ This tells you there is a written assignment.

This tells you that this is the key to the answers for the self-tests.

UNIT 4



PERIODIC CLASSIFICATION OF ELEMENTS

Unit Outcomes

Dear learner, after completing the unit you will be able to

- Explain the historical development of the periodic classification of the elements
- Describe the periodic classification of the elements.
- Develop the skills of correlating the electron configuration of elements with the periodicity of the elements, predicting the trends of periodic properties of elements in the periodic table.
- Develop skills of classifications based on patterns in chemistry.
- Demonstrate scientific inquiry skills: observing, inferring, predicting, classifying, comparing and contrasting, making models, communicating, measuring, asking questions, interpreting illustrations, drawing conclusion, applying concepts and problem solving.

Start Up Activity

Conduct the following activity before you start reading study notes and compare it with your understanding after you get done.

How can you use a table of repeating events to predict the next events? ***Table 2.1*** shows a familiar table of repeating properties. What is it?

Of course, it is a calendar of month of Nehassie, 2013 E.C and it is missing some information. You can determine what is missing and fill in the blanks.

1. Look at the calendar again. The calendar has columns of days, Monday through Sunday. The calendar also has horizontal rows. They are its weeks.
2. Examine the information surrounding each empty spot. Can you tell what information is needed in each empty spot?
3. Fill in the missing information.

For example: Conclude and Apply

1. One day in column 4 is marked X, and a day in column 5 is marked Y. What dates belong to these positions?
2. Column 3 does not have a name. What is the correct name of this column?
3. What dates are included in the third row of the table?
4. What day would the 25th of the previous month (month of Hamle, 2013 E.C) have been? What row of this table would it appear in?
5. How do you relate this periodicity with the periodic classification of elements?

Table 4.1 Periodicity in month.

Mon	Tue		Thu	Fri	Sat	Su
					1	2
3	4	5	6	7	8	9
10	11	12	X	y	15	16
17	18	29	20	21	22	23
24	25	26	27	28	29	30

Unit Content

In this unit you are going to cover the following sections

Section 4.1: Historical Development of Periodic Classification of the Elements

Section 4.2: The Modern Periodic Table

Section 4.3: Classification of the Elements

Section 4.4: The Representative Elements

Section 4.5: The Major Trends in the Periodic Table (Atomic Size, Ionization Energy, Electron Affinity and Electro Negativity)

The Required Study Time

- Starting from the first day of your enrollment, you have a maximum of 10 months to complete the grade 9 chemistry modules, but it also depends on the pace at which you proceed. Grade 9 chemistry consists of five units and this is the unit 4 of this course. You are expected to finish studying this unit in eight weeks.

Unit Learning Strategies

As a distance learner, you are expected to learn by yourself by carefully studying the study notes, doing the activities, self-test exercises, self-assessment exercises and assignment. As a student, you are responsible for mastering the lessons, completing all the activities given in this unit, self-test exercises, unit self-assessment questions, and the assignment. You are also responsible for checking your work carefully, noting areas in which you need to improve and motivating yourself to succeed. Using the checklist evaluate yourself whether you mastered the key points of the unit or not and try to revise on the areas of your weaknesses. Hereunder are the strategies you should follow in your study. Read the components of each section and do the instructions to get the maximum knowledge benefits out of it.

Each section in this unit consists of the following components:

Section Title

The section title simply provides you with the title of the section.

Section Overview

Each section begins with the outline of what you are going to learn in each of the sections. Please read it carefully to get the impression of what you gain from the section before you explore the main contents of the section.

Section Learning Outcomes

Each lesson has learning outcome which are competencies you should have to accomplish by the end of the lesson. You should make sure that you have achieved them before moving to the next section.

Study Notes and Important Points

The main body of the lesson is consists of the content that you need to learn. It contains definition of terms, explanations, diagrams, and fully completed examples. You need to read and understand the definition of terms and the detailed explanations of the concepts and ideas.

Resources

All the content is provided directly within the course. When you are having difficulty with something in this unit, contact your learning partner or tutor/marker, which is there to help you.

Learning Activities

The learning activities that are given in this unit include questions that you should complete in order to help you practice what you have learned in each section, prepare you to complete your assignment and the examination successfully, achieve the learning outcomes of the sections. Make sure that you have completed all the learning activities. Once you have completed the learning activities, you should check your answers with the answer key provided at the end of the unit. You will not submit the completed learning activities to the Distance Learning Unit.

Definition

The summary of definitions of terms is provided in each section following the resources. This is to help you get all the definitions of terms being summarized. Make sure that you can define all terms orally and through writing and clearly understood their meaning.

Self-Test Exercise

The self-test exercises provided to help you achieve the learning outcomes of each section. There are five self-test exercises in this unit. You must do all the exercises and check answers found at the end of the unit, afterwards. You do not need to give the solutions to your tutor/marker.

Checklist

In order to help you audit whether or not you have done everything that is required of you and mastered the learning outcomes of each section, there is a checklist in each section that you need to go through it. Fill the checklist being honest. Revise the things you did not perform very well until you are clear about them.

Unit Summary

The unit ends with a brief review of the main points of what you just learned. You are advised to read it so that you can get the summary of the main points of the unit.

Self Assessment

In this unit there is one self-assessment at the end of the unit containing 19 questions with three parts namely true or false, multiple choice and filling in blank spaces. The questions are categorized into three levels; basic, intermediate and Advanced level. To help you succeed in your examination, you will have an opportunity to complete the self-assessment questions. The self-assessment is similar to the actual examination you will be writing. The answer key enables you to check your answers. This will give you the confidence you need to do well on your examination.

Assignment

In this unit there is one assignment with 20 questions. The assignment has to be done individually. For the numerical questions, you have to show all the necessary steps in your calculations. For questions that need explanation, you are expected to give a detailed explanation. Follow all the given instructions for each group of questions. You need to submit the report for your tutor/marker for correction.

Answer Keys

At the end of the unit there are answers for section level activity questions, self-test, and self-assessment exercises. Do not refer to these answers before you do all the activity questions, the self-test and self-assessment exercise questions by yourself. Do not also leave without checking out your answers after you have done all the exercises.

Section 4.1: Historical Development of Periodic Classification of the Elements

- Classification is often one of the important tasks in a scientific study. It becomes particularly important when we need to study particular aspects of a collection of large numbers of things. Dealing with the different aspects each individual member of the collection may become difficult and be unnecessary. Thus, look for a way of classifying or grouping them into smaller numbers of classes. After reducing the size of our sample to a manageable minimum, we can then concentrate only on the behaviour of each class since each individual in a class share some common characteristics.

Learning Competencies

At the end of this section, you should be able to

- Describe periodicity.
- State Mendeleev's Periodic law.

Two important examples of classification can be given. First Carolus Linus' classification system in which the huge number of living things (plants and animals) are grouped into a very small number of classes. Without the use of sucha classification system, the science of biology would not be possible.

The second example of the periodic system of classification of chemical elements, used in chemistry and the other physical sciences. In this system of classification, the one hundred and eleven known elements are placed in a few categories, in a chart known as the periodic table.

The periodic table consists of seven horizontal rows of elements called **periods** each of which are divided in to two eighteen vertical columns, known as groups. The periodic table show that there is a repetitive pattern, or periodicity of the properties of the elements. Elements belonging to the same group show very close similarity in their properties. Elements in the same period show gradual variations in their properties as we move through the period. Thus, the behaviour of the elements can be inferred from their positions in the periodic table. The discovery of the periodic table was not an accident. Nor was it the work of individual chemist. Rather it is the outcome of the painstaking efforts of several chemists.

This unit begins by looking of four of these individuals. This review will be followed by an analysis of the modern periodic table. Finally you will meet some examples of both the use and the imperfections of the periodic table.

4.1.1. Dobereiner's Triads

In 1829, German chemist Dobereiner was able to identify several groups of three elements that showed similarity in physical and chemical properties. He observed that in the set of three elements having similar properties (called triads), the atomic weight of the middle element is the arithmetic mean of the atomic weights of other two elements. Some examples of Dobereiner's triads are as follows (***Table 4.2***):

Table 4.2: Exampes of Dobereiner's triads

	Element	At. Weight	Mean weight of first and last element
(i)	Li Na K	7 23 39	$\frac{7+39}{2}=23$
(ii)	Ca Sr Ba	40 88 137	$\frac{40+137}{2}=88.5$
(iii)	Cl Br I	35.5 80 127	$\frac{35.5+127}{2}=81.25$
(iv)	S Se Te	32 79 127.6	$\frac{32+127.6}{2}=79.8$

Only few such triads were available at that time and day by day as many more elements were discovered, the rule could no longer be generalized.

4.1.2. Newland's Law of Octaves.

An English chemist Alexander Newlands made the next attempt at classification of elements. He arranged the 56 elements known then in increasing order of their atomic weight and observed that, "the properties of every eighth element are similar to that of the first one". He compared this relationship to the first octave in music with eight notes and called it Newlands' law of octaves. The elements were arranged as follows (***Table 4.3***):

Table 4.3: Newlands' law of octaves

Sa	Re	Ga	Ma	Pa	Dha	Ni	Sa
Li	Be	B	C	N	O	F	Na
Na	Mg	Al	Si	P	S	Cl	K
K	Ca						

The limitations of this classification of elements were that:

i. The inert gases were not discovered till then.

ii. Beyond Ca, this repetition was not observed.

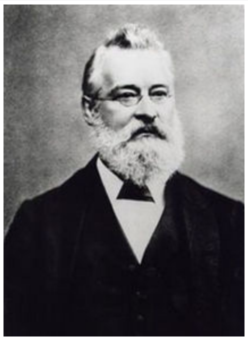


Figure 4.1: John Alexander Reina Newlands was an English chemist who worked on the development of the periodic table. He noticed that elemental properties repeated every seventh (or multiple of seven) elements, as musical notes repeat every eighth note.

According to Newlands there will be a recurrence of properties when elements are arranged in order of increasing atomic masses

Activity 4.1

Time allocated: 5 minutes

The relationship to the first octave in music with eight notes is observed in musical instrument. That means the first sound repeats itself on the eighth (Do Ri Mi Fa So La Si Do). Can you explain how the Ethiopian cultural music does the same. To answer this question you can consult an expert in playing musical instruments.

4.1.3. Mendleev's Classification of the Elements

The earliest version of the current form of periodic table was presented simultaneously by Dmitri Mendeleev of Russia and Lothar Meyer of Germany. Both the scientists arranged the elements in order of increasing atomic weights and observed that elements with similar properties (in families) appeared at regular intervals.

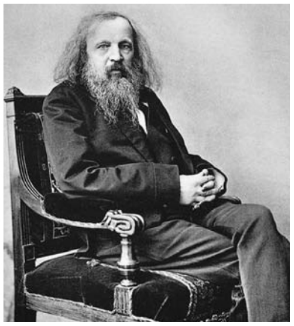


Figure 4.2: Dmitri Ivanovich Mendeleev (1834–1907) Mendeleev constructed a periodic table as part of his effort to systematize chemistry. He received many international honors for his work, but his reception in czarist Russia was mixed. He had pushed for political reforms and made many enemies as a result.

4.1.4.Mendeleev's Periodic Law

Mendeleev's periodic law stated that the physical and chemical properties of elements are the periodic function of their atomic weights.

Mendeleev was able to show that the chemical properties of the elements exhibit repetitive patterns when arranged in order of increasing atomic masses. His idea may be illustrated as follows.

Let us begin

In 1871, Mendeleev published a short periodic table which:

1. Consisted of only 63 elements. Inert gases were not included as these were not discovered at that time.
2. These elements were arranged in seven horizontal rows called as periods and eight vertical columns called as groups.
3. Some vacant sites were specified for undiscovered elements and their properties predicted. These were found true and verified when these elements were discovered later.

In the periodic table, so constructed, the elements in the same families (e.g., lithium, sodium, potassium) were arranged in vertical columns designated as Groups I, II, III, IV, V, VI, VII, VIII. The horizontal rows were referred to as series.

Mendeleev's periodic table was later modified after the discovery of inert gases and several other elements. The inert gases were placed in new Group 0. Each long period was divided into two series, named as odd and even depending on the serial number. The first seven elements formed the even series and the last seven elements formed the odd series (not including the inert gases). The vertical Groups I to VII were further divided into two subgroups A and B to accommodate elements with difference in properties. The elements of even series in the long periods were placed in subgroup A while the elements of odd series were placed in the B subgroup. The Group 0 was not split further and in Group VIII three sets containing three elements each were placed.

Exercise 4.1

Give appropriate answers for the following questions.

1. Which group of elements was missing from Mendeleev's periodic table as compared to the modern periodic table?
2. Why does the first period contain only two elements?
3. To which group and period do the following elements belong?
   a. carbon
   b. neon
   c. aluminium
   d. potassium
   e. calcium
   f. sulphur

Table 4.4: Modern version of Mendeleev's short periodic table

Group / Period	0	I A	I B	II A	II B	III A	III B	IV A	IV B	V A	V B	VI A	VI B	VII A	VII B	VIII
1	He 2	H 1														
2	Ne 10	Li 3		Be 4			B 5		C 6		N 7		O 8		F 9	
3	Ar 18	Na 11		Mg 12			Al 13		Si 14		P 15		S 16		Cl 17	
4		K 19		Ca 20		Sc 21		Ti 22		V 23		Cr 24		Mn 25		Fe 26, Co 27, Ni 28
	Kr 18		Cu 29		Zn 30		Ga 31		Ge 32		As 33		Se 34		Br 35	
5		Rb 37		Sr 38		Y 39		Zr 40		Nb 41		Mo 42		Tc 43		Ru 44, Rh 45, Pd 46
	Xe 54		Ag 47		Cd 48		In 49		Sn 50		Sb 51		Te 52		I 53	
6		Cs 55		Ba 20		La* 57-71		Hf 72		Ta 73		W 74		Re 75		Os 76, Ir 77, Pt 78
	Rn 86		Au 79		Hg 80		Tl 81		Pb 82		Bi 83		Po 84		At 85	
7		Fr 87		Ra 88		Ac** 85-103										

The Rare Earths

*Lanthanide series (6th period)	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
**Actinide series (7th period)	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

4.1.5. Periodicity

Important characteristics of modern version of Mendeleev's short periodic table are listed as follows:

I. Horizontal Rows or Periods

1. First period consists of 2 elements and is known as very short period.
2. Second period consists of 8 elements and is known as first short period.
3. Third period consists of 8 elements and is known as second short period.
4. Fourth period consists of 18 elements and is known as first long period.
5. Fifth period consists of 18 elements and is known as second long period.
6. Sixth period consists of 32 elements.
7. Seventh period also consists of 32 elements and both period six and period seven are known as very loang periods.

II. Vertical Columns or Groups

1. Group IA elements are called as alkali metals (expect H).
2. Group IIA elements are called as alkaline earth metals.
3. Group VB elements are called as pnicogens.
4. Group VIB elements are called as chalcogens.
5. Group VIIB elements are called as halogens.
6. Group 0 elements are called as noble (inert) gases.

The merits of Mendeleev's periodic table are listed as follows:

1. The study of properties of elements became more systematic and easier.
2. There are several vacant positions from which the guidance of discovery of new elements was found.

The demerits of Mendeleev's periodic table are given as follows:

Some of the elements are wrongly placed though their atomic weights are larger compared to the next one. For example,

(i)	Ar:40	(ii)	Te: 127.6	(iii)	Co: 58.9
	K:39		I: 126.9		Ni: 58.6

Moseley's Periodic Law

Moseley showed that the positions of the elements in the periodic table depend on atomic numbes (number of protons in the atom) not on atomic masses. He demonstrated that if elements are arranged in order of increasing atomic numbers, all the anomalies in Mendeleev's periodic table will be removed. For example, the interchange between potassium and argon, as discussed above, will not occur, since potassium (atomic number 19) comes after argon (atomic number 18) in the order. Moseley formulated a law known as modern periodic law, which may be stated as follows:

Properties of the elements are periodic functions of their atomic numbers.

On the basis of this law, several forms of the periodic table have been constructed. One form known as short form will be described in the next section.

Checklist 4.1

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
describe periodicity?		
state Mendeleev's periodic law?		
state and describe the basis of		
obereiner's triads?		
newlands Octaves?		
Mendeleev's periodic table?		
define the following terms		
period?		
groups?		
periodic table?		
demonstrate the existence of some patterns in the properties of the elements using examples?		

Self-Test Exercise 4.1

I. Shortly answer the following questions?

1. What was the basis of Dobereiner's classification of the elements into a number of triads?
2. What was the basis of Newland's classification of the elements into "Octaves"?
3. Can you suggest Newland's contribution to Mendeleev's discovery?
4. Mention at least two ways by which Mendeleev's classification of the elements differs from that of Newland's.
5. Cite three examples to show that Mendeleev's periodic law was not correctly stated.
6. State Moseley's periodic law.
7. Why are the chemical properties of the elements related to their atomic numbers, rather than their atomic mass?

II. Matching the following scientist with their contributions.

	Column A	Column B
1.	Dobereiner	A. Periods and groups of elements
2.	Newland	B. Modern periodic Law
3.	Menderleev	C. Triads of Elements
4.	Moseley	D. Octave of elements

Section 4.2: The Modern Periodic Table

- The modern periodic table shows the arrangements of elements in order of increasing atomic numbers. In the table, elements are arranged in horizontal row, called **periods**, and vertical column, called **groups**. The place each element occupies in the table corresponds to the electron configuration of its atom. This section reviews the structure of the periodic table of the elements, and the relationships between the position an element occupies in the table and the electron configuration of its atom.

Learning Competencies

At the end of this section, you should be able to

- State modern periodic law.
- Describe period.
- Describe group.
- Classify the periods into short, long and incomplete periods.

s-Block (ns) | d-Block, (n–1)d | p-Block (np) (nonmetals)

Group / Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18 •
	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA ••
1 (1s)	H 1		Transition metals (d-block)														H 1	He† 2
2 (2s, 2p)	Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
3 (3s, 3p)	Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
4 (4s, 3d, 4p)	K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
5 (5s, 4d, 5p)	Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
6 (6s, 4f, 5d, 6p)	Cs 55	Ba 56	La* 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
7 (7s, 5f, 6d, 7p)	Fr 87	Ra 88	Ac** 89	Unq 104 (Rf)	Unq 105 (Db)	Unq 106 (Sg)	Unq 107 (Bh)	Unq 108 (Hs)	Unq 109 (Mt)	Unq 110	Unq 111	Unq 112						

• New convention
•• Earlier convention

Post transition metals

The Rate Earths, (n–2)f

f-block															
*Lanthanides series(4f) Period :6	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
**Actinide series(5f) Period :7	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

Figure 4.3: The long form of modern periodic table and its segments.

4.2.1.The Periodic Law

According to the modern periodic law, the physical and chemical properties of the elements are the periodic functions of their atomic number. The long form of periodic table based upon the modern periodic law is depicted in ***Figure 4.3***. Note that the arrangement of A and B subgroups is different from that in the modified form of Mendeleev's periodic table. The left and right corners of the table are assigned as sub-groups A and the middle of the periodic table is assigned as subgroups B.

Elements in the same group have similar chemical properties.

4.2.2.Groups and Periods

Many different forms of the periodic table have been published since Mendeleev's time. Today, the long form of the periodic table, which is called the modern periodic table, is commonly in use. It is based on the modern periodic law. In the modern periodic table, elements are arranged in periods and groups.

? What are the basis for classifying the elements into groups and periods? What are the similarities and differences in the electron configuration of S nand Cl?

Periods: The horizontal rows of elements in the periodic table are called periods or series. Elements in a period are arranged in increasing order of their atomic numbers from left to right. There are 7 periods in the modern periodic table, and each period is represented by an Arabic numeral: 1, 2 . . . and 7.

Elements in the same period have the same number of shells.

Periods 1, 2, and 3 are called short periods while periods 4, 5, 6 and 7 are known as long periods.

Period 1 contains only 2 elements, hydrogen and helium. Period 2 and period 3 contain 8 elements each.

Period 4 and period 5 contain 18 elements each. Period 6 and 7, the longest period, has 32 elements. Period 7 elements are radioactive and/or an artificial elements.

Except for the first period, all periods start with an alkali metal and ends with a noble gas.

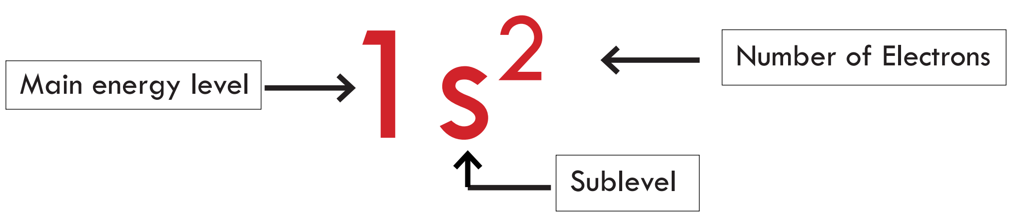
Table 4.5: The number of elements in a given period and the orbitals being filled.

Period number	Orbitals occupied	Number of elements
1	1s	2
2	2s, 2p	8
3	3s, 3p	8
4	4s, 3d, 4p	18
		and so on

The position of an element in a given period can be determined by the number of shells occupied with its electrons. Accordingly, the number of shell is equal to the number of period to which the element belongs.

4.2.3. Electron Configurations

The arrangement of electrons in an atom is known as the electron configuration of the atom. Because atoms of different elements have different numbers of electrons, a distinct electronic configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest energy arrangement of the electrons in an atom is called the ground state electron configuration. A few simple rules, combined with the quantum number relationships discussed below, allow us to determine these ground state electron configurations. The quantum mechanical model is designated by the following notation: a coefficient which shows the main energy level, a letter that denotes the sublevel that an electron occupies, and a superscript that shows the number of electrons in that particular sublevel. The designation is explained as follows



For example, the electron configuration of lithium (${}^{5}_{3}Li$) is: $1s^2 2s^1$. This indicates that there are 2 electrons in the first s-sublevel and 1 electron in the second s-sublevel. The configuration for sodium (${}^{23}_{11}Na$) atom is: $1s^2 2s^2 2p^6 3s^1$. This indicate that there are 2 electrons in the first s-sublevel, 2 electrons in the second s-sublevel, 6 electrons in the second p-sublevel, and 1 electron in the third s-sublevel.

The lowest energy sublevel is always the 1s sublevel, which consists of one orbital. The single electron of the hydrogen atom will occupy the 1s orbital when the atom is in its ground state. As we proceed to atoms with multiple electrons, those electrons are added to the next lowest sublevel: 2s, 2p, 3s, and so on. The order of filling orbitals with electrons is according to the Aufbau principle, which states that an electron occupies orbitals in order from lowest energy to highest. The Aufbau (German for building up, construction) principle is sometimes referred to as the "building up" principle. It is worth noting that in reality, atoms are not built by adding protons and electrons one at a time, and that this method is merely an aid to understand the end result.

As seen in the figure above, the energies of the sublevels in different principal energy levels eventually begin to overlap. After the 3p sublevel, it would seem logical that the 3d sublevel should be the next lowest in energy. However, the 4s sublevel is slightly lower in energy than the 3d sublevel and thus fills first. Following the filling of the 3d sublevel is the 4p, then the 5s and the 4d. Note that the 4f sublevel does not fill until just after the 6s sublevel. ***Figure 4.4*** is a useful and simple aid for keeping track of the order of fill of the atomic sublevels.

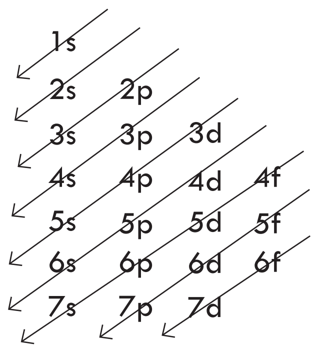


Figure 4.4 The diagonal rule for electron filling order

Example 4.1

Electronic configuration of $^{23}_{11}Na = 1s^2 2s^2 2p^6 3s^1$ (2, 8, 1). Sodium has 3 main shells. Hence, sodium is found in period 3.

Activity 4.2

Time allocated: 5 minutes

By looking example above write the electron configuration of 7F and 17Cl. What are the similarities and differences in the electron configuration of F and Cl?

Groups or families: are the vertical columns of elements in the periodic table. There are 18 columns or groups in the modern periodic table. Group numbers are usually designated with the numbers 1 to 8 each followed by the letter A or B.

These are:

IA . . . . VIIIA Main groups (A groups)
IB . . . . VIIIB Sub groups (B groups)

Elements in a given group have the same number of outermost shell electrons. Elements in the same group have similar chemical properties. For the main group elements the group number equals the number of valence electrons.

Because elements in the same group have similar properties, they may be identified by family names. For example, the group IA elements are known as **alkali metals**. Those in group IIA are called the **alkaline earth elements**. Every member in these two groups is a reactive metal and undergoes chemical reactions similar to those of the other member of its respective groups. The elements at the right most corner of the periodic table are known as noble gases. When we reach period 4, we encounter the rows of elements that divide the elements of group IIA and IIIA. These group of elements are identified by the labels IB to VIIIB, and are called the **transition metals**.

? Dear learner! Can you suggest why these elements are known with the name "transition metals"?

The transition metals form a bridge between the chemically active metals of group II A and the much less reactive group III A metals. There is a block of two long rows of elements placed below the main table in ***Figure 4.5***. The elements found in this block are termed inner transition elements.

? Dear learner! Why are these elements so named?

The block is drawn there, simply to save space. The upper row of this block that follows after element lanthanum is known as lanthanide series of elements, whereas the lower row, following actinium is the actinide series.

Checklist 4.2

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
describe period and group?		
state modern periodic law?		
classify the periods into short, long and incomplete periods?		
define the following terms		
periods of elements?		
group of elements?		
periodic table?		
describe the basis of classifying the elements in the modern periodic table?		
list down the rules of Aufbau principle?		
explain the orders of filling orbitals with electrons?		
predict the possible electronic structure of an atom?		
write the electronic structures of atoms following Aufbau procedure?		
identify incorrect electronic structures?		

Self-Test Exercise 4.2

Multiple choice Questions

1. The elements in group VIIA are known as:
   A. Transition metals
   B. Halogens
   C. Alkali metals
   D. Alkaline earth metals
   E. Noble gases
2. The elements in group IIA are known as
   A. Transition metals
   B. Halogens
   C. Alkali metals
   D. Alkaline earth metals
   E. Noble gases
3. The alkali metals are found in _______ of the periodic table
   A. Group IA
   B. Group IIA
   C. Group IIIA
   D. Period 7
   E. Period 1
4. Which of the following elements is a transition metal?
   A. Sr
   B. Pb
   C. As
   D. Fe
   E. H
5. Which of the following elements is a transition metal?
   A. Iron
   B. Tin
   C. Sodium
   D. Sulphur
   E. Calcium

Section 4.3: Classification of the Elements

? What is the relationship between the positions of the elements in the periodic table and the electron configurations of their atoms?

- The organization of the periodic table will be clear once we knew the electron configurations of their atoms. All atoms in a given period have valence electrons in the same shell, and each new period begins with the occupation of a new shell. Therefore, the principal quantum number of an element's valence shell is equal to its period number. In other words, the label of a period in which an element is found is the same as the principal quantum number of the last orbital that its electrons occupy. For example, magnesium is a period 3 element because if we write the electron configuration of Mg atom as $1s^22s^22p^63s^2$, the n value for the last electron (or 3s electrons) is 3. The fact that each period corresponds to the n-value also explains why different periods have different lengths.

Learning Competencies

At the end of this section, you should be able to:

- Explain the relationship between the electronic configuration and the structure of the modern Periodic table.
- Describe the three classes of the elements in the modern Periodic table.
- Give group names for the main group elements.
- Predict the period and group of an element from its atomic number.
- Tell the block and group of an element from its electronic configuration.

> All the elements in the same period have the same value of n, the principal quantum number, for its last electrons. Conversely, all elements in the same period have the same core structures.

Elements of period 2 have "He"-like core, and those of period 3 have "Ne"-like core. In periods 1 to 3, atoms of the elements of each period differ only in their numbers of valence electrons. Elements in the same group differ in their value of principal quantum number, n. they have similar valence shell structure, i.e., a structure with a different core.

> The similarity in electron configurations of the elements in the same group are the reasons for their similar properties.

The group numbers tell us how many valence electrons are present in the atoms of the elements. For example, all the elements of group IA have one electron in their valence shells. Those in group VIIA have seven valence electrons.

All the main group (or representative, that will be discussed in the next section) elements have their valence electrons in the s- and p-sublevels. The transition metals have valence electrons in the s- and d-sublevels. All of these elements have 2 electrons in the s-sublevel, except Cr and Cu. Therefore the major difference between the transition metals is due to the numbers of d-electrons that they have.

All the elements of the noble gas have 8 electrons in their valence shell, except helium. Atoms of inner transition elements have incomplete f orbitals to which electrons are being added. The lanthanide differ only in their numbers of f electrons. All of them have already filled s and p orbitals. This is the reason for their chemical similarity. In actinide, the 5f orbitals are being filled. ***Figure 4.5*** shows a periodic table that indicates which orbitals are being filled by electrons as we move through that various regions of the table and ***Figure 4.6*** the detain electronic configuration on the last orbitals are given for each elements. Elements in the left most columns, i.e., groups IA and IIA, which have valence electrons in s-orbital are called s-block elements. Those in the 6 columns from Groups IIIA to VIIIA that have valence electrons in the p-orbitals are known as the p-block elements. The transition metals haved-valence electrons and hence are termed d-block elements. The inner transition elements are known as f-block elements.

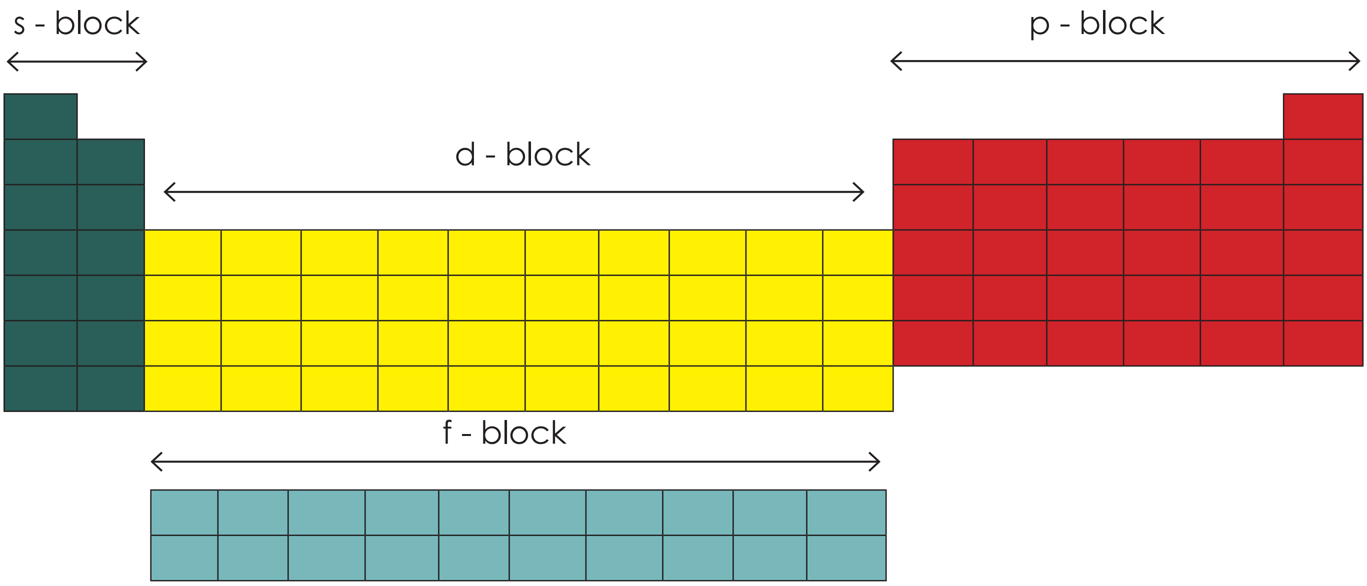


Figure 4.5 The form of the periodic table that shows the various blocks of elements: s-block ( ), p-block ( ), d-block ( ), and f-block ( )

Checklist - 4.3

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
locate in which period and group of the periodic table a given element if found by observing its electron configuration?		
predict the properties of an element based on its position on the periodic table?		
explain the relationship the position an element occupies on the periodic table and the electron configuration of its atom?		
locate the regions on the periodic table where the following elements in the bracket can be found (K, Ca, Fe, H, F, Hg, U, and Pb)?		
distinguish between the various families of elements based on their general chemical properties?		

Self-Test Exercise 4.3

1. The general electron configuration of atoms of all elements in group VA is:
   A. ns^2np^6
   B. ns^2np^5
   C. ns^2np^4
   D. ns^2np^3
   E. ns^2np^1
2. Which of the following is the general electron configuration for the outer most electrons of elements in the alkaline earth group?
   A. ns^1
   B. ns^2
   C. ns^2np^4
   D. ns^2np^5
   E. $ns^2np^6(n^{-1})d^6$
3. An element with the general electron configuration for its outer most electrons of ns^2np^1 would be in which group?

A. IIA
B. IIIA
C. IVA
D. VA
E. VIIA

4. In what group of the periodic table is the element with the following electron configuration $[Ar]4s^2 3d^{10} 4p^3$ found?
A. IA
B. IIA
C. IIIA
D. IVA
E. VA

5. Where do you classify the element with the following electron configuration $[Kr]5s^2 4d^7$?
A. A representative
B. A transition element
C. A non-metal
D. an actinide element
E. a noble gas

6. Which two of the following electron configurations represent elements that would have similar chemical properties?
1. $1s^2 2s^2 2p^4$ 2. $1s^2 2s^2 2p^5$ 3. $[Ar]4s^2 3d^{10} 4p^3$ 4. $[Ar]4s^2 3d^{10} 4p^5$
A. 1, 3
B. 1, 2
C. 1, 4
D. 2, 4
E. 2, 3

Section 4.4: The Representative Elements

- Referring to ***Figure 4.6***, the representative elements (also called main group elements) are the elements in Groups IA through VIIA, all of which have incompletely filled s or p subshells of the highest principal quantum number. With the exception of helium, the noble gases (the Group VIIIA elements) all have a completely filled p subshell. (The electron configurations are $1s^2$ for helium and ns^2np^6 for the other noble gases, in which n is the principal quantum number for the outermost shell.)

Learning Competencies

- At the end of this section, you should be able to explain the representative element related with their electronic configuration.

A clear pattern emerges when we examine the electron configurations of the elements in a particular group. The electron configurations for Groups IA and IIA elements are shown in ***Figure 4.6***. We see that all members of the Group IA alkali metals have similar outer electron configurations; each has a noble gas core and an ns^1 configuration of the outer electron. Similarly, the Group IIA alkaline earth metals have a noble gas core and an ns^2 configuration of the outer electrons. The outer electrons of an atom, which are those involved in chemical bonding, are often called **the valence electrons**. Having the same number of valence electrons accounts for similarities in chemical behavior among the elements within each of these groups. This observation holds true also for the halogens (the Group VIIA elements), which have outer electron configurations of ns^2np^5 and exhibit very similar properties. We must be careful, however, in predicting properties for Groups

IIIA through VIA. For example, the elements in Group IVA all have the same outer electron configuration, ns^2np^4, but there is much variation in chemical properties among these elements: Carbon is a nonmetal, silicon and germanium are metalloids, and tin and lead are metals.

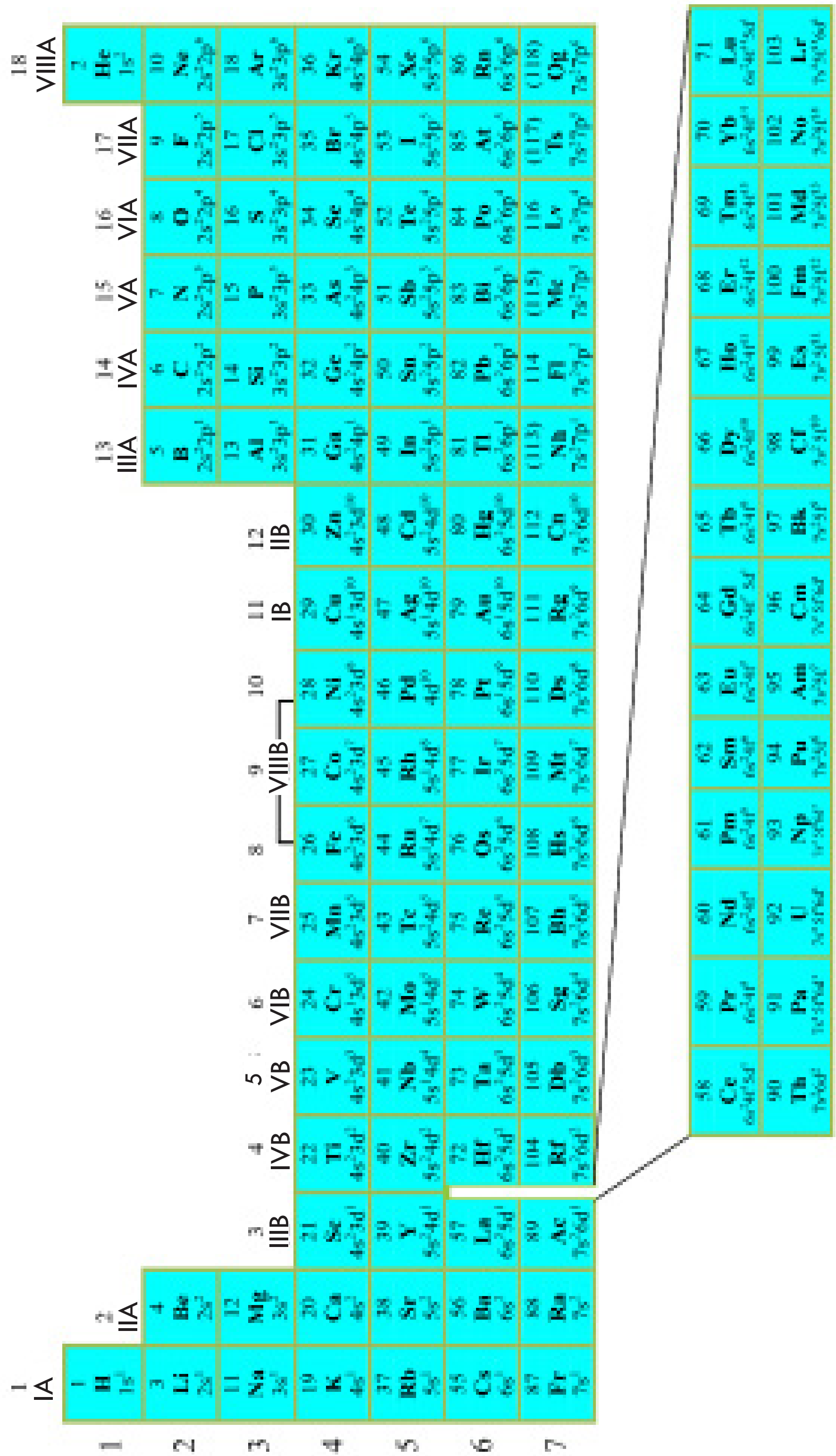


Figure 4.6 The ground-state electron configurations of the elements. For simplicity, only the configurations of the outer electrons are shown.

Checklist - 4.4

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
explain the representative element related with their electronic configuration?		
distinguish representative elements from others?		
predict properties for groups IA , IIA and Groups IIIA through VIA?		

Self-Test Exercise 4.4

1. Where do you classify the element with the following electron configuration [Ne]$3s^2 3p1$
   A. A representative
   B. A transition element
   C. A non-metal
   D. An actinide element
   E. A noble gas
2. The representative elements are those with unfilled energy levels in which the "last electron" was added to:
   A. An s-orbital
   B. An s- or a p-orbital
   C. A d-orbital
   D. A p- or a d-orbital
   E. An f-orbital

Section 4.5: The Major Trends in the Periodic Table (Atomic Size, Ionization Energy, Electron Affinity and Electro Negativity)

- In the section 4.3 you have seen that if elements are arranged in order of increasing atomic numbers, there will be some kind of repetition of their properties after some intervals of elements (or atomic numbers). In this section you will exanmine some examples of properties which show periodicity.

Learning Competencies

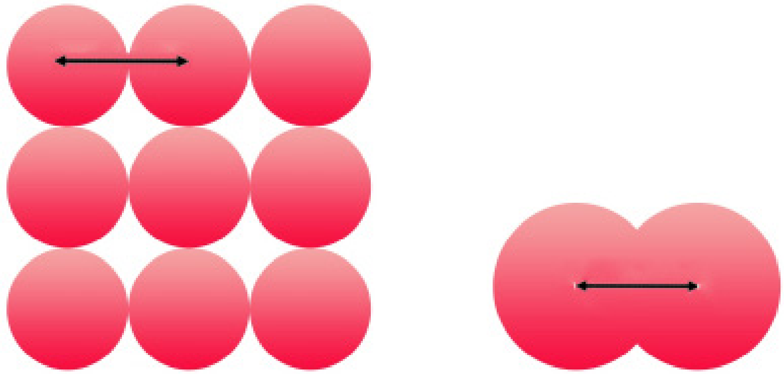
At the end of this section, you should be able to describe the four major trends (atomic size, ionization energy, electron affinity and electro negativity) in the periodic table.

As we have seen, the electron configurations of the elements show a periodic variation with increasing atomic number. Consequently, there are also periodic variations in physical and chemical behavior. In this section, we will examine some physical properties of elements that are in the same group or period and additional properties that influence the chemical behavior of the elements.

4.5.1. Atomic Radius

A number of physical properties, including density, melting point, and boiling point, are related to the sizes of atoms, but atomic size is difficult to define. When we must be even more specific, we define the size of an atom in terms of its atomic radius, which is one-half the distance between the two nuclei in two adjacent metal atoms.

For atoms linked together to form an extensive three-dimensional network, atomic radius is simply one-half the distance between the nuclei in two neighbouring atoms [***Figure 4.7(a)***]. For elements that exist as simple diatomic molecules, the atomic radius is one-half the distance between the nuclei of the two atoms in a particular molecule [***Figure 4.7(b)***].



(a) (b)

Figure 4.7: (a) In metals such as beryllium, the atomic radius is defined as one-half the distance between the centers of two adjacent atoms. (b) For elements that exist as diatomic molecules, such as iodine, the radius of the atom is defined as one-half the distance between the centers of the atoms in the molecule.

Dear learner! Now let us examine the trends of atomic size on the periodic table. If we start at Li and move down the group, from the top to the bottom of the periodic table, there will be an increase the atomic radii of the elements. H is the smallest and Fr is the largest (see ***Figure 4.8***) although the nuclear charge increases, the outer most electrons begin in a new shells, and this effect out weighs the increased nuclear attraction. Similar behaviour is found for other groups in the periodic table.

Atomic radii generally increase from top to bottom down a group.

Now let's turn to the trend as we move across a period. Suppose we start at Li and move to the right to F. We find a decrease of size from Li to F (see ***Figure 4.8***).

? Can you suggest the largest and the smallest members of period 2 elements? Why?

If you suggest that lithium is the largest and fluorine is the smallest you are right. This may be explained by the following fact.

As new electrons are added in the same shell of the atom, they will be about as close to nucleus as the other electrons in the same shell. But since the nuclear charge also increases by one unit for each added electron, the electrons will be drawn with a greater attraction. This results in a reduction in size. But when an octet of electrons (that means eight electrons) are added to the shell, there will be some type of relaxation in period 2 indeed the noble gas Neon (Ne) is larger than the fluorine (F).

Atomic radii generally decrease as we go from left to right across the periodic table.

In explaining the reason for such a trend in atomic radius the terms effective nuclear charge and shielding effect are used. The effective nuclear charge is the actual amount of positive (nuclear) charge experienced by an electron in a multi-electron atom. The term "effective" is used because the shielding effect of negatively charged electrons prevent higher energy electrons from experiencing the full nuclear charge of the nucleus due to the repelling effect of inner layer. The effective nuclear charge experienced by an electron is also called the core charge. In chemistry, the shielding effect sometimes referred to as atomic shielding or electron shielding describes the attraction between an electron and the nucleus in any atom with more than one electron. The shielding effect can be defined as a reduction in the effective nuclear charge on the electron cloud, due to a difference in the attraction forces on the electrons in the atom.

Figure 4.8 shows the atomic radii of many elements according to their positions in the periodic table. Periodic trends are clearly evident. In studying the trends, bear in mind that the atomic radius is determined to a large extent by the strength of the attraction between the nucleus and the outer-shell electrons. The larger the effective nuclear charge, the stronger the hold of the nucleus on these electrons, and the smaller the atomic radius. Consider the second-period elements from Li to F, for example. Moving from left to right, we find that the number of electrons in the inner shell ($1s^2$) remains constant while the nuclear charge increases. The electrons that are added to counterbalance the increasing nuclear charge are ineffective in shielding one another. Consequently, the effective nuclear charge increases steadily while the principal quantum number remains constant (n=2). For example, the outer 2s electron in lithium is shielded from the nucleus (which has three protons) by the two 1s electrons. As an approximation, we assume that the shielding effect of the two 1s electrons is to cancel two positive charges in the nucleus. Thus, the 2s electron only feels the attraction of one proton in the nucleus; the effective nuclear charge is +1. In beryllium ($1s^2 2s^2$), each of the 2s electrons is shielded by the inner two 1s electrons, which cancel two of the four positive charges in the nucleus. Because the 2s electrons do not shield each other as effectively, the net result is that the effective nuclear charge of each 2s electron is greater than +1. Thus, as the effective nuclear charge increases, the atomic radius decreases steadily from lithium to fluorine.

Within a group of elements we find that atomic radius increases with increasing atomic number. For the alkali metals in Group 1A, the outermost electron resides in the ns orbital. Because orbital size increases with the increasing principal quantum number n, the size of the metal atoms increases from Li to Cs even though the effective nuclear charge also increases. We can apply the same reasoning to the elements in other groups.

Activity 4.4

Time allocated: 10 minutes

Discuss why:

1. All the elements of a group have similar chemical properties.
2. All the elements of a period have different chemical properties.
3. The atomic radii of three elements A, B and C of a period of the periodic table are 186 pm, 104 pm and 143 pm respectively. Giving a reason, arrange these elements in the increasing order of atomic numbers in the period.

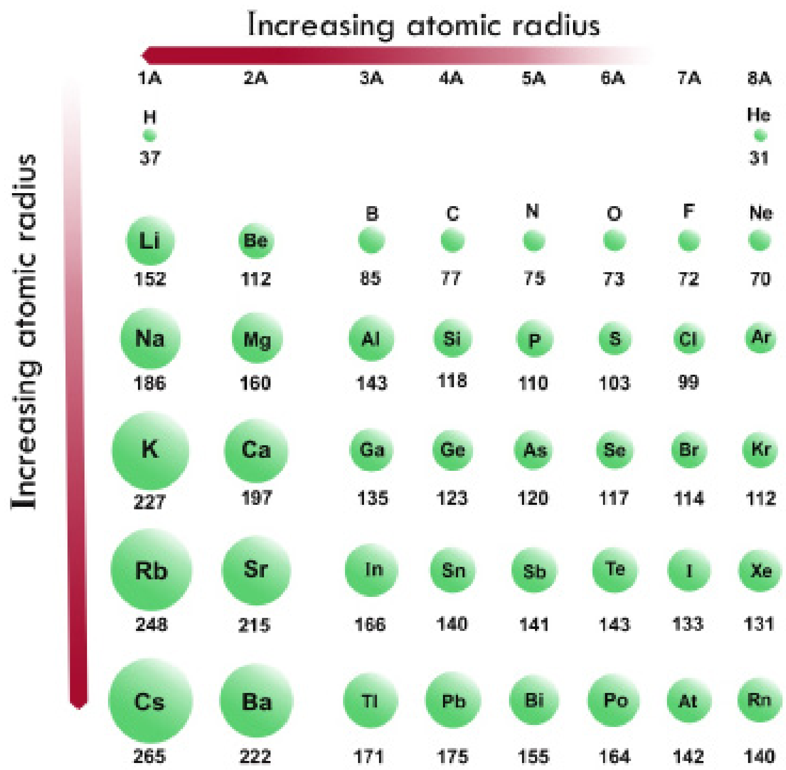


Figure 4.8: Atomic radii (in picometers) o representative elements according to their positions in the periodic table. Note that there is no general agreement on the size of atomic radii. We focus only on the trends in atomic radii, not on their precise values.

Example 4.3:

Referring to a periodic table, arrange the following atoms in order of increasing atomic radius: P, Si, N.

Strategy What are the trends in atomic radii in a periodic group and in a particular period? Which of the preceding elements are in the same group? in the same period?

Solution: From ***Figure 4.5*** we see that N and P are in the same group (Group 5A). Therefore, the radius of N is smaller than that of P (atomic radius increases as we go down a group). Both Si and P are in the third period, and Si is to the left of P. Therefore, the radius of P is smaller than that of Si (atomic radius decreases as we move from left to right across a period). Thus, the order of increasing radius is $N < P < Si$.

Exercise 4.2:

Arrange the following atoms in order of decreasing radius: C, Li, Be and give explanation to your arrangement.

4.5.2. Ionization Energy

Ionization is the process of removing electrons from an atom and ionization energy required to remove an electron from an atom in the gas phase. Electrons in atom are under the attractive force of a nucleus. Therefore, to remove an electron we must supply energy that is sufficient to overcome the attractive force.

Ionization energy is the quantity of energy that an isolated, gaseous atom in the ground electronic state must absorb to discharge an electron, resulting in a cation.

$$H(g) \rightarrow H^+(g) + e^- \quad (1)$$

This energy is usually expressed in kJ/mol, or the amount of energy it takes for all the atoms in a mole to lose one electron each.

The ionization energy is the minimum amount of energy that we must supply to remove an electron from an atom. The amount of energy that we supply depends on a number of factors. The size of the atom is one important factor. Larger atoms have electrons at a greater distance from the nucleus than smaller atoms. Since attractive forces diminish with increasing distances it requires less energy to remove an electron from a larger atom. Note that during ionization the last electron is removed first. Since the size of atoms increase down a group, ionization energies decrease as we move from top to bottom a given group.

? Dear learner! Do you recall the Aufbau process?

During the Aufbau process, electrons are added to orbitals in order of increasing energy. The first electron goes to the orbital that is closest to the nucleus. After this orbital is completely filled, we pass to filling the next orbitals successively. The electron last added goes into the orbital farthest from the nucleus. The ionization process may be thought as the opposite of the Aufbau process. When we remove the electrons first we remove the electron last added and then successively pass to the others. The amount of energy that we supply to remove the first electron from an atom, **first ionization energy (IE_1)** is the smallest. To remove the second electron more amount of energy, **second ionization energy (IE_2)**, is needed. Still greater amount of energy is needed to remove the third, fourth electron and so on.

For a given atom
$IE_1 < IE_2 < IE_3 < \ldots\ IE_n$

? Dear learner! Can you explain the reasons for the above order of ionization energies?

The second factor that determines the magnitude of the ionization energy is the effective nuclear charge (Zeff). Zeff increases from left to right across a period, but is nearly the same for all the elements in the same group. This explains why ionization energy increases from left to right through a period.

Now let us look at the values of the first ionization energies for period 2 elements in ***Table 4.6***. If we examine these values, we see some irregularities in the order of the ionization energies of these elements. It is true that the ionization energies increase from left to right, for most of the elements. However, we notice two elements (B and O) that break the regularity, i.e., decreases, rather than increases, were observed as we move from Be to B, and from N to O. the two factors we discussed above, atomic size and effective nuclear charge, cannot account for the irregularities. Therefore, we must look for other factors.

Now let's examine the type of orbitals from which the electrons are removed. First, let's compare the electron configuration of the pairs of elements.

Be: $1s^22s^2$ N: $1s^22s^22p^3$
B: $1s^22s^22p^1$ O: $1s^22s^22p^4$

We know that it is the last electron that is first removed. During ionization process:

Be: $1s^22s^2 \rightarrow Be^+: 1s^22s^1$ N: $1s^22s^22p^3 \rightarrow N^+: 1s^22s^22p^2$
B: $1s^22s^22p^1 \rightarrow B+: 1s^22s^2$ O: $1s^22s^22p^4 \rightarrow O^+: 1s^22s^22p^3$

In the case of Be and B, the electrons are removed from different types of orbitals. We compare the relative distances from the nucleus of the orbitals of a shell, the s-orbital is closest to the nucleus followed by p, d, and f, in that order. The electron ins-orbitals are closer to the nucleus than those in p-orbital, and hence, require greater ionization energies. This explains why the first ionization energy of B is smaller than that of Be.

Now lets us turn to the case of N and O. Both the electron of N and O are removed from the same type of orbitals. However, the electron of N is removed from one of the p orbitals, each of which contains one electron. When the degenerate orbitals are all filled either one or two electrons, the electron distribution becomes so uniform or symmetrical, so that removal of an electron does not take place readily. The oxygen atom attain this symmetry when it lose one of its electrons to become O^+. Hence, removal of the first electron from O requires less energy than that of N.

Table 4.6: Showing the increasing trend of ionization energy in kJ/mol (exception in case of Boron) from left to right in the periodic table

Li	Be	B	C	N	O	F
520	899	800	1086	1402	1314	1680

When considering an initially neutral atom, expelling the first electron will require less energy than expelling the second, the second will require less energy than the third, and so on. Each successive electron requires more energy to be released. This is because after the first electron is lost, the overall charge of the atom becomes positive, and the negative forces of the electron will be attracted to the positive charge of the newly formed ion. The more electrons that are lost, the more positive this ion will be, the harder it is to separate the electrons from the atom.

In general, the further away an electron is from the nucleus, the easier it is for it to be expelled. In other words, ionization energy is a function of atomic radius; the larger the radius, the smaller the amount of energy required to remove the electron from the outer most orbital. For example, it would be far easier to take electrons away from the larger element of Ca (Calcium) than it would be from one where the electrons are held tighter to the nucleus, like Cl (Chlorine).

In a chemical reaction, understanding ionization energy is important in order to understand the behavior of whether various atoms make covalent or ionic bonds with each other. For instance, the ionization energy of Sodium (alkali metal) is 496 kJ/mol whereas Chlorine's first ionization energy is 1251.1 kJ/mol. Due to this difference in their ionization energy, when they chemically combine they make an ionic bond. Elements that reside close to each other in the periodic table or elements that do not have much of a difference in ionization energy make polar covalent or covalent bonds. For example, carbon and oxygen make CO_2 (Carbon dioxide) reside close to each other on a periodic table; they, therefore, form a covalent bond. Carbon and chlorine make CCl_4 (Carbon tetrachloride) another molecule that is covalently bonded. The different types of chemical bonds will be discussed in Unit 5.

As described above, ionization energies are dependent upon the atomic radius. Since going from right to left on the periodic table, the atomic radius increases, and the ionization energy increases from left to right in the periods and up the groups. Exceptions to this trend is observed for alkaline earth metals (Group IIA) and nitrogen group elements (Group VA). Typically, Group IIA elements have ionization energy greater than Group IIIA elements and group VA elements have greater ionization energy than group VIA elements. Groups IIA and VA have completely and half-filled electronic configuration respectively, thus, it requires more energy to remove an electron from completely filled orbitals than incompletely filled orbitals.

Activity 4.5

Time allocated: 5 minutes

Rank each set of the following elements in order of decreasing ionization energy and explain the trend in ionization energy of the elements down a group.

a. Ca, Sr, Mg, Be

b. K, Li, Rb, Na

c. Cl, F, I, Br

Alkali metals (Group IA) have small ionization energies, especially when compared to halogens or Group 8A (see ***Table 4.6***). In addition to the radius (distance between nucleus and the electrons in outermost orbital), the number of electrons between the nucleus and the electron(s) you're looking at in the outermost shell have an effect on the ionization energy as well. This effect, where the full positive charge of the nucleus is not felt by outer electrons due to the negative charges of inner electrons partially canceling out the positive charge, is called shielding. The more electrons shielding the outer electron shell from the nucleus, the less energy required to expel an electron from said atom. The higher the shielding effect the lower the ionization energy (see ***Table 4.7***). It is because of the shielding effect that the ionization energy decreases from top to bottom within a group. From this trend, Cesium is said to have the lowest ionization energy and Fluorine is said to have the highest ionization energy (with the exception of Helium and Neon).

Table 4.7: showing decreasing trend of ionization energies (kJ/mol) from top to bottom (Cs is the exception in the first group)

Li 520
Na 496
K 419
Rb 408
Cs 376
Fr 398

Example 4.5

a. Which atom should have a smaller first ionization energy: oxygen or sulfur?

b. (Which atom should have a higher second ionization energy: lithium or beryllium?

Strategy (a) First ionization energy decreases as we go down a group because the outermost electron is farther away from the nucleus and feels less attraction. (b) Removal of the outermost electron requires less energy if it is shielded by a filled inner shell.

Solution (a) Oxygen and sulfur are members of Group VIA. They have the same valence electron configuration (ns^2np^4), but the 3p electron in sulfur is farther from the nucleus and experiences less nuclear attraction than the 2p electron in oxygen. Thus, we predict that sulfur should have a smaller first ionization energy.

(b) The electron configurations of Li and Be are $1s^22s^1$ and $1s^22s^2$, respectively. The second ionization energy is the minimum energy required to remove an electron from a gaseous unipositive ion in its ground state. For the second ionization process we write

$$Li^+(g) \rightarrow Li^{2+}(g) + e^-$$
$$1s^2 \qquad\qquad 1s^1$$
$$Be^+(g) \rightarrow Be^{2+}(g) + e^-$$
$$1s^22s^1 \qquad\qquad 1s^2$$

Because 1s electrons shield 2s electrons much more effectively than they shield each other, we predict that it should be easier to remove a 2s electron from Be^+ than to remove a 1s electron from Li^+

Exercise 4.3:
(a) Which of the following atoms should have a larger first ionization energy: N or P?
(b) Which of the following atoms should have a smaller second ionization energy: Na or Mg?

4.5.3. Electron Affinity

Electron affinity is defined as the change in energy (in kJ/mole) of a neutral atom (in the gaseous phase) when an electron is added to the atom to form a negative ion. In other words, the neutral atom's likelihood of gaining an electron.

Activity 4.6

Time allocated: 10 minutes

1. Why does the electron affinity of Cl is higher than that of F?
2. Explain why noble gases have extremely low (almost zero) electron affinities?
3. Explain why halogens have the highest electron affinities?

? When the first electron is added, energy is usually released, but the addition of the second electron may require energy. Why?

Electron affinity is the energy released or absorbed when an electron is added to an atom in the gas phase. Nonmetals have the tendency to accept electrons.

The first electron goes to a neutral atom. After receiving an electron the atom become a negatively charged ion. Therefore, the second electron goes to an anion. There will be a repulsion between the second electron and the anion. To overcome this repulsive force energy must be supplied. For example, when the first electron is added to an atom of oxygen 141 KJ/mol of energy is released. But the addition of the second electron requires of 844 KJ/mol of energy.

Elements towards the right top corner of the periodic table have the highest electron affinities.

Electron affinity generally increased from left to right across a period and decrease from top to bottom through a group on the periodic table.

Energy of an atom is defined when the atom loses or gains energy through chemical reactions that cause the loss or gain of electrons. A chemical reaction that releases energy is called an exothermic reaction and a chemical reaction that absorbs energy is called an endothermic reaction. Energy from an exothermic reaction is negative, thus energy is given a negative sign; whereas, energy from an endothermic reaction is positive and energy is given a positive sign. An example that demonstrates both processes is when a person drops a book. When he or she lifts a book, he or she gives potential energy to the book (energy absorbed). However, once he or she drops the book, the potential energy converts itself to kinetic energy and comes in the form of sound once it hits the ground (energy released).

When an electron is added to a neutral atom (i.e., first electron affinity) energy is released; thus, the first electron affinities are **negative**. However, more energy is required to add an electron to a negative ion (i.e., second electron affinity) which overwhelms any the release of energy from the electron attachment process and hence, second electron affinities are **positive**.

First Electron Affinity (negative energy because energy released):

$$X(g) + e^- \rightarrow X^-(g) \quad (1)$$

Second Electron Affinity (positive energy because energy needed is more than gained):

$$X^-(g) + e^- \rightarrow X^{2-}(g) \quad (2)$$

Ionization energies are always concerned with the formation of positive ions. Electron affinities are the negative ion equivalent, and their use is almost always confined to elements in Groups 6 and 7 of the Periodic Table. The first electron affinity is the energy released when 1 mole of gaseous atoms each acquire an electron to form 1 mole of gaseous -1 ions. It is the energy released (per mole of X) when this change happens. First electron affinities have negative values. For example, the first electron affinity of chlorine is -349 kJ mol-1. By convention, the negative sign shows a release of energy.

When an electron is added to a metal element, energy is needed to gain that electron (endothermic reaction). Metals have a less likely chance to gain electrons because it is easier to lose their valance electrons and form cations. It is easier to lose their valence electrons because metals' nuclei do not have a strong pull on their valence electrons. Thus, metals are known to have lower electron affinities.

Example 4.6 Group IA Electron Affinities
This trend of lower electron affinities for metals is described by the Group I metals:

- Lithium (Li): -60 KJ mol^{-1}
- Sodium (Na): -53 KJ mol^{-1}
- Potassium (K): -48 KJ mol^{-1}
- Rubidium (Rb): -47 KJ mol^{-1}
- Cesium (Cs): -45 KJ mol^{-1}

Notice that electron affinity decreases down the group.

When nonmetals gain electrons, the energy change is usually negative because they give off energy to form an anion (exothermic process); thus, the electron affinity will be negative. Nonmetals have a greater electron affinity than metals because of their atomic structures: first, nonmetals have more valence electrons than metals do, thus it is easier for the nonmetals to gain electrons to fulfill a stable octet and secondly, the valence electron shell is closer to the nucleus, thus it is harder to remove an electron and it easier to attract electrons from other elements (especially metals). Thus, nonmetals have a higher electron affinity than metals, meaning they are more likely to gain electrons than atoms with a lower electron affinity.

Example 4.6: Group 7A Electron Affinities
For example, nonmetals like the elements in the halogens series in Group 7A have a higher electron affinity than the metals. This trend is described as below. Notice the negative sign for the electron affinity which shows that energy is released.

- Fluorine (F) -328 kJ mol^{-1}
- Chlorine (Cl) -349 kJ mol^{-1}
- Bromine (Br) -324 kJ mol^{-1}
- Iodine (I) -295 kJ mol^{-1}

Notice that electron affinity decreases down the group, but increases up with the period. As the name suggests, electron affinity is the ability of an atom to accept an electron. Electron affinity is a quantitative measurement of the energy change that occurs when an electron is added to a neutral gas atom. The more negative the electron affinity value, the higher an atom's affinity for electrons. ***Figure 4.9*** shows the electron affinity trend.

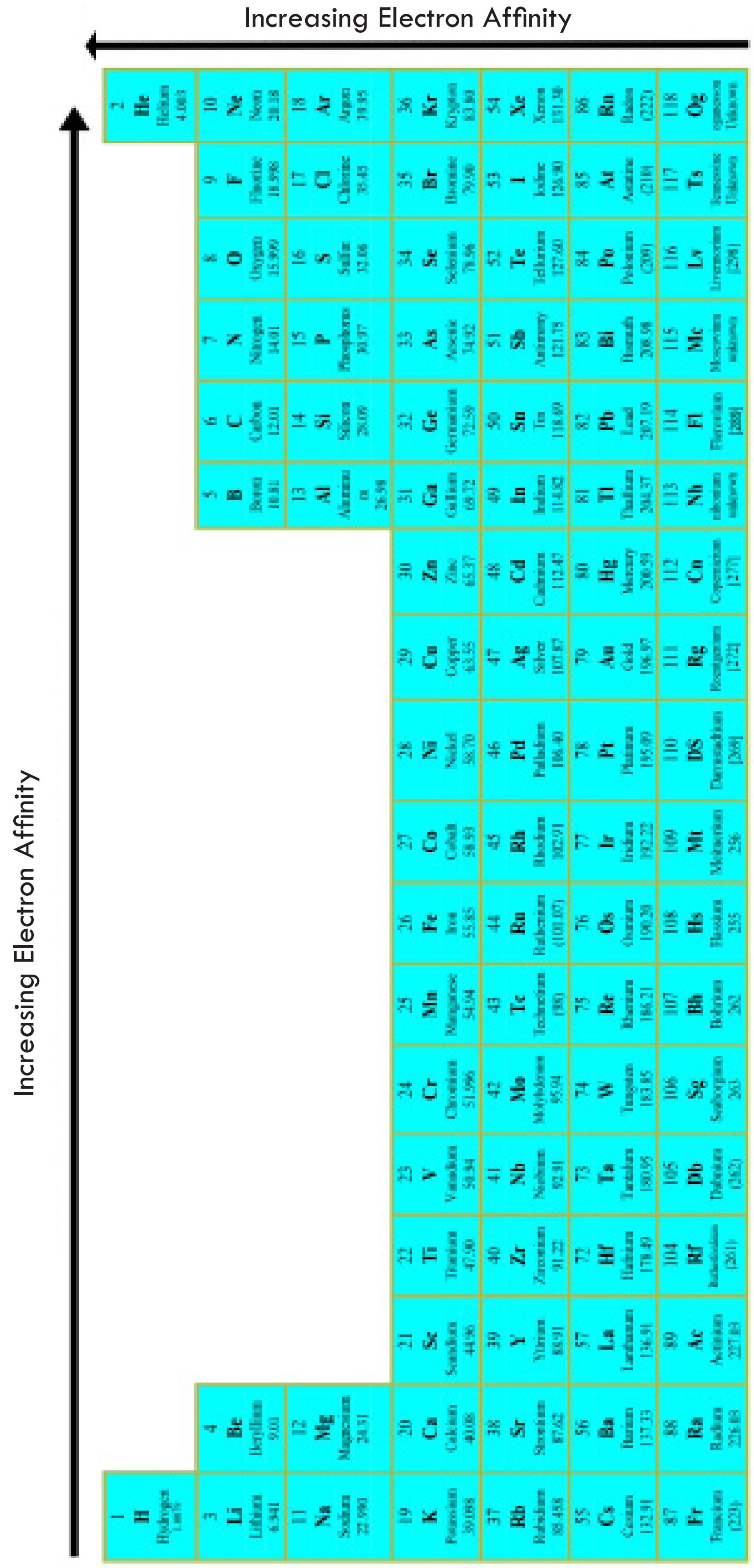


Figure 4.9 Periodic Table showing Electron Affinity Trend.

Figure 4.9: Periodic Table showing Electron Affinity Trend

4.5.4. Electronegativity

Electronegativity is the measure of the tendency of an atom to attract electrons to itself. It is a dimensionless property because it is only a tendency. It basically indicates the net result of the tendencies of atoms in different elements to attract the bond-forming electron pairs. We measure electronegativity on several scales. The most commonly used scale was designed by Linus Pauling. According to this scale, fluorine is the most electronegative element with a value of 4.0 and cesium is the least electronegative element with a value of 0.7.

The tendency of an atom in a molecule to attract the shared pair of electrons towards itself is known as electronegativity.

As we move across a period from left to right the nuclear charge increases and the atomic size decreases, therefore the value of electronegativity increases across a period in the modern periodic table. There is an increase in the atomic number as we move down the group in the modern periodic table. The nuclear charge also increases but the effect of the increase in nuclear charge is overcome by the addition of one shell. Hence, the value of electronegativity decreases as we move down the group. For example, in the halogen group as we move down the group from fluorine to astatine the electronegativity value decreases.

Electronegativity generally increase from left to right across a period and decreases from top to bottom down a group on the periodic table.

Activity 4.7

Time allocated: 10 minutes

Arrange each set of the given elements in order of decreasing electronegativity and explain the observed trend.

A. Ba, Mg, Be, Ca

B. C, Pb, Ge, Si

C. Cl, F, I, Br

It is a general observation that metals show a lower value of electronegativity as compared to the non-metals. Therefore, metals are electropositive and non-metals are electronegative in nature. The elements in period two differ in properties from their respective group elements due to the small size and higher value of electronegativity. The elements in the second period show resemblance to the elements of the next group in period three. This happens due to a small difference in their electronegativities. This leads to the formation of a diagonal relationship (see section 4.2.2).

Those elements requiring only a few electrons to complete their valence shells, and having the least quantity of inner electron shells between the positive nucleus and the valence electrons, are the most electronegative. The most electronegative of all elements are fluorine. Its electronegativity is 4.0. Metals have electronegativities less than 2.0. The least electronegative elements are cesium (Cs), with electronegativity values of 0.7. Therefore, Fluorine is the most electronegative element and cesium is the least electronegative element.

? Dear learner! Which one is more electronegative an alkali metal or a halogen element?

Checklist - 4.5

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
define the following properties:		
Atomic radius?		
Ionization energy?		
Electron affinity?		
Electronegativity		
describe the variations in the above properties in the periodic table when going through periods and groups of the periodic table?		
use the periodic table to arrange the elements in orders of increasing or decreasing atomic size, ionization energy, electron affinity, and electronegativity?		
locate where irregularities in the trends of variations the periodic properties occur?		
explain the reasons why the irregularities in these trends occur		

Self-Test Exercise 4.5

Use the periodic table to answer the following questions.

1. Which atom of the pair is larger? Be or Fe, Mg or Ca, Cl or I, Ti or W, F or Ne, S or I, Cr or Ni?
2. By looking into ***Figure 4.8***, put the following atoms in order of increasing size: Sr, Na, P, Xe, Be.
3. Put the following elements in order of increasing electron affinity: S, P, Al, Si, and Cl.
4. Arrange the following elements in order of increasing ionization energy: K, Na, Li, Cs, and Rb.
5. Arrange the following elements in decreasing order of electronegativity: O, B, N, C, and F.

II. Multiple choice Questions: Choose the best answer among the given alternatives.

6. Which element will have the largest ionization energy?

A. Cs C. K E. As

B. Ga D. Bi

7. Which element will have the greatest ionization energy?

A. C C. Ge D. Sn

B. Si E. Pb

8. Which element will have the lowest ionization energy?

A. Cl C. Be E. As

B. Na D. K

9. Which element will have the lowest ionization energy?

A. Li C. Be E. Rb
B. Na D. K

10. Which element will show an unusually high jump in ionization energy between Z_3 and Z_4?

A. Na C. Al E. P
B. Mg D. Si

11. Which element has the largest electron affinity?

A. Mg C. Si E. S
B. Al D. P

12. Which element has the largest electron affinity?

A. K C. As E. I
B. Br D. Ar

13. Which element has the largest electronegativity?

A. Na C. Ga E. Sb
B. As D. Cs

14. Which element has the largest electronegativity?

A. Li C. P E. Ge
B. Cs D. As

15. Which element has the largest electronegativity?

A. Se C. K
B. Sb D. Ga E. Fe

16. Which element has the least electronegativity?

A. Sr C. Ni
B. V D. P

Unit Summary

Nineteenth-century chemists developed the periodic table by arranging elements in the increasing order of their atomic masses. Discrepancies in early versions of the periodic table were resolved by arranging the elements in order of their atomic numbers.

The elements can be arranged in rows and columns by atomic number to form the periodic table. Elements in a given group (column) have similar properties. (A period is a row in the periodic table.) Elements on the left and at the center of the table are metals; those on the right are nonmetals.

Electron configuration determines the properties of an element. The modern periodic table classifies the elements according to their atomic numbers, and thus also by their electron configurations. The configuration of the valence electrons directly affects the properties of the atoms of the representative elements.

Atomic radius varies periodically with the arrangement of the elements in the periodic table. It decreases from left to right and increases from top to bottom.

Ionization energy is a measure of the tendency of an atom to resist the loss of an electron. The higher the ionization energy, the stronger the attraction between the nucleus and an electron.

Electron affinity is a measure of the tendency of an atom to gain an electron. The more positive the electron affinity, the greater the tendency for the atom to gain an electron. Metals usually have low ionization energies, and nonmetals usually have high electron affinities.

Noble gases are very stable because their outer ns and np subshells are completely filled.

The metals among the representative elements (in Groups 1A, 2A, and 3A) tend to lose electrons until their cations become isoelectronic with the noble gases that precede them in the periodic table.

The nonmetals in Groups 5A, 6A, and 7A tend to accept electrons until their anions become isoelectronic with the noble gases that follow them in the periodic table.

The tendency of an atom in a molecule to attract the shared pair of electrons towards itself is known as electronegativity. It is a dimensionless property because it is only a tendency. It basically indicates the net result of the tendencies of atoms in different elements to attract the bond-forming electron pairs.

Self-Assessment Exercises

Part I: Basic level questions.

Identify whether each of the following statements is true or false. Give your reasons when you consider a statement to be false.

1. The modern periodic law was proposed by Mosley.
2. Elements across a period have consecutive atomic numbers.
3. All elements with high ionization energy also have high electron affinity.
4. As the atomic number of elements increases in the periodic table, their atomic radius also increases.
5. Transition metals are found in four periods. Each corresponds to the filling of valence electrons in the 3d, 4d, 5d, and 6d orbitals.

PART II: Intermediate level questions.

Given below are multiple choice questions. Choose the best answer from the given alternatives.

6. The periodic law states that
   a. similar properties recur periodically when elements are arranged according to increasing atomic number
   b. similar properties recur periodically when elements are arranged according to increasing atomic weight
   c. similar properties are everywhere on the periodic table
   d. elements in the same period have same characteristics
7. Which element is most similar to Sodium
   a. Potassium
   b. Al uminum

c. Oxygen
d. Calcium

8. Which element is most similar to Calcium?
   a. Carbon
   c. Strontium
   b. Oxygen
   d. Iodine
9. Who were the two chemists that came up with the periodic law?
   a. John Dalton and Michael Faraday
   b. Dmitri Mendeleev and Lothar Meyer
   c. Michael Faraday and Lothar Meyer
   d. John Dalton and Dmitri Mendeleev
10. The statement that is not true about electron affinity is
   a. It causes energy to be released
   b. It causes energy to be absorbed
   c. It is expressed in electron volts
   d. It involves formation of an anion
11. Which of the following is Dobereiner's triad?
   a. Ne, Ca, Na
   c. Li, Na, K
   b. H_2, N_2, O_2
   d. Na, Br, Ar
12. Write the period number, group number and block of the element having atomic number 42.
   a. 5, 5, d
   c. 5, 2, d
   b. 5, 6, d
   d. 5, 15, p
13. X, Y and Z are three consecutive elements. X on addition of one electron and Y on addition of two electrons become isoelectronic with element Z. Which of the following is the correct property of Element Y?
   a. Atomic number of Y is higher than atomic number of Z.
   b. Atomic number of Y is higher than atomic number of X.
   c. Element Y is placed in periodic table at left side of element X if both are in same period.
   d. None of these

To which group, period, and sublevel block do the following elements belong?

	Group	Period	Block type
Magnesium	______	______	______
Phosphorous	______	______	______
Krypton	______	______	______
Manganese	______	______	______
Gold	______	______	______
Potassium	______	______	______

Assignment for Submission

Direction: Attempt all questions. Give a detailed explanation.

1. Briefly describe the significance of Mendeleev's periodic table.
2. What is Moseley's contribution to the modern periodic table?
3. Describe the general layout of a modern periodic table.

4. What is the most important relationship among elements in the same group in the periodic table?
5. A neutral atom of a certain element has 17 electrons. Without consulting a periodic table, (a) write the ground-state electron configuration of the element, (b) classify the element.
6. Group these electron configurations in pairs that would represent similar chemical properties of their atoms:
   a. $1s^22s^22p^63s^2$
   b. $1s^22s^22p^3$
   c. $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$
   d. $1s^22s^2$
   e. $1s^22s^22p^6$
   f. $1s^22s^22p^63s^23p^3$
7. Without referring to a periodic table, write the electron configurations of elements with these atomic numbers: (a) 9, (b) 20, (c) 26, (d) 33. Classify the elements.
8. Specify in what group of the periodic table each of these elements is found:
   a. [Ne]3s1,
   b. $[Ne]3s^23p^3$,
   c. $[Ne]3s^23p^6$, (d) $[Ar]4s^23d^8$.
9. How does atomic radius change as we move (a) from left to right across the period and (b) from top to bottom in a group?
10. On the basis of their positions in the periodic table, select the atom with the larger atomic radius in each of these pairs: (a) Na, Cs; (b) Be, Ba; (c) N, Sb; (d) F, Br; (e) Ne, Xe.
11. Arrange the following atoms in order of decreasing atomic radius: Na, Al, P, Cl, Mg.
12. Define ionization energy. Ionization energy is usually measured in the gaseous state. Why? Why is the second ionization energy always greater than the first ionization energy for any element?
13. (a) Define electron affinity. Electron affinity is usually measured with atoms in the gaseous state. Why? (b) Ionization energy is always a positive quantity, whereas electron affinity may be either positive or negative. Explain.
14. Arrange the elements in each of these groups in order of increasing electron affinity: (a) Li, Na, K; (b) F, Cl, Br, I.
15. Define Electronegativity. How does electronegativity vary as we move (a) from left to right across the period and (b) from top to bottom in a group?

Answer Key to Exercises

Self-Test Exercise I

I. Short answer question

1. There are two main bases: (1) the chemical reactivity between the elements and (2) the patterns in their atomic masses.
2. The basis of Newland's system of classification is the existence of some kind of connection between the chemical properties of the elements. That means, the chemical properties of the elements repeat themselves ater some intervals when elements are listed in increasing orders of atomic masses.
3. Newland's contribution is his introduction of the idea of octaves, which are similar to what Mendeleev named as periods of elements.

4. The biggest and the most important difference between Newland's and Mendeleev's classification systems is that Mendeleev left gaps where the patterns of atomic masses break. A second difference may be the fact thatMendeleev's system was introduced later in time and hence included larger number of elements.
5. K comes before Ar, I before Te, Ni before Co.
6. Statement Mosle's periodic law: properties of the elements are periodic functions of their atomic numbers.
7. As we will see in the next unit, chemical bonding, they are the electrons that take part in bond formation. Since the atomic number is the number of electrons contained by a neutral atom, there can be some relationship between the atomic numbers of the elements and their chemical properties, the mass number does not have any direct relationship with the chemical properties, though it may influence physical properties.

II. Matching

1. C 2. D 3. A 4. B

Self-Test Exercise 2

I. Multiple choice

1. B 2. D 3. A 4. D 5. A

Self-Test Exercise 3

I. Multiple choice item

1. D 2. B 3. B 4. E 5. B 6. D

Self-Test Exercise 4

I. Multiple choice item

1. A 2. B

Self-Test Exercise 5

I. Short Answer

1. Be, Ca, I, W, Ne, I, Cr
2. P (110) < Be(112) < Xe(131) < Na(186) < Sr(215)
3. Al < Si < P < S < Cl
4. Cs < Rb < K < Na < Li
5. F > O > N > C > B

II. Multiple choice

1. E	4. E	7. B	10. B
2. A	5. C	8. B	11. A
3. D	6. C	9. C	

Answer to Self-Assessment Exercises

Part I: True or False Items

1. True	3. False	5. True
2. True	4. False	

Part II Multiple choice

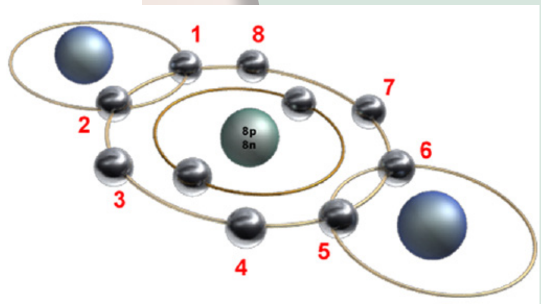
6. A 7. A 8. C 9. B

10. B   11. C   12. B   13. C

Part III:

Group	Period	Block type
______IIA______	______3______	_______s____
______VA______	______3______	_______p____
______VIIIA___	______4______	_______p____
______VIIB____	______4______	_______d____
______IB______	______6______	_______d____
______IA______	______4______	_______s____

UNIT 5



CHEMICAL BONDING

Introduction

? Dear learner, did you remember what you learned about the structure of an atom in Unit 3? What have you learned about the relationship between periods or groups and the properties of the elements in Unit 4?

In this unit, the main focus will be on the association of the electrons and the properties of the elements in chemical bond formation. Elements are distinguishable from each other due to their electrons. Because each element has a distinct number of electrons, this determines its chemical properties as well as the extent of its reactivity. In chemical bonding, only valence electrons are involved. In addition to this, the unit deals with the way the different types of chemical bonds are formed, the properties of compounds formed in each type of chemical bonding, how to write the formulas used to represent molecules or compounds formed in each type of chemical bonding, and the concept and origin of the polarity of molecules. The contents are divided into five sections and presented sequentially. You should therefore study each section and get a good understanding of its contents.

Unit Outcomes

Dear learner, at the end of this unit, you should be able to

- Discuss the formation of ionic, covalent and metallic bonds.
- Explain the general properties of substances containing ionic, covalent and metallic bonds.
- Develop the skills of drawing the electron dot or Lewis structures for simple ionic and covalent compounds.
- Describe the origin of polarity within molecules.

Unit Contents

Section 5.1: Chemical Bonding
Section 5.2: Ionic Bonding
Section 5.3: Covalent Bonding
Section 5.4: Metallic Bonding

The Required Study Time

From the date of your registration, you have a maximum of 10 months to complete the course, but the pace at which you proceed is up to you. This course has a total of five units. Proportionally dividing the period allotment for the teacher guided- lesson (13 periods out of 61) into 10 months, the time you need to study this unit is nine weeks.

Section 5.1: Chemical Bonding

Overview

? What can you say about elements and their representation, valence number, meaning of chemical formula, writing the formula of binary compounds, writing chemical formulas, polyatomic ions, and writing and balancing chemical equations, from your Grade 7, Unit 2 lessons? Did you remember the meanings of mono, di, and polyatomic molecules of homoatomic nature, from your Grade 8 Unit 2 lessons? Do you know how to determine the valence electrons of an element? What is the relationship between valence electrons and group numbers in the periodic table of the elements?

All the contents you have studied so far will help you understand why elements combine to form molecules or compounds. You, however, have not studied why atoms combine or react to form molecules and compounds thus far. In this section, therefore, you are going to study the meaning of chemical bonding and the reason why atoms form chemical bonds.

Learning Outcomes

Dear learner, when you have finished studying this section, you should be able to

- Define chemical bonding.
- Describe why atoms form chemical bonds.

Study Notes and Important Points

Dear learner, study the following notes and important points before you do the activity and self-test exercise questions. Make sure that you are clear on the meanings of the concepts, terms, and definitions. After doing so, check whether you achieved the learning outcomes indicated above.

Let me begin presenting the contents by asking you some basic questions.

? Based on your previous science lessons, can you define the terms atom, element, molecule, and compound? Most of the elements are not found in their atomic form. Do you know why this is so? What binds atoms together to form compounds or molecules? Do you know why atoms combine to form compounds or molecules?

The material world in which we live is changing from time to time. This is because nature is associated with innumerable chemical and physical changes. Chemical changes occur due to changes in the composition of the elements from which the materials are made. We call these chemical changes chemical reactions.

As you have learned in your Grade 7 science lesson, a chemical reaction involves a change in the molecular composition of a substance.

It has been established that **an atom** is the smallest particle of matter that takes part in a chemical reaction. **A molecule** is the smallest constituent particle of a substance that has an independent existence, and which represents the properties of the respective elements or compounds.

? Dear learner, do you know what a chemical bond is? Why do atoms form a chemical bond?

Most of the elements in nature have been found to exist in a combined state. This is because the atoms of some elements are unstable by themselves. There are, however, some atoms that exist uncombined with other elements and are highly stable. Except for a few elements, the atoms of most of the elements have a characteristic tendency to combine and form molecules or compounds. Those atoms may belong to the same or different elements. In the molecules, the atoms that form the molecules are held together by attractive forces. **A chemical bond** is, therefore, the force that holds atoms together to form molecules or compounds.

The American chemist Gilbert Lewis (1875-1946) observed on the number of valence electrons in atoms when various types of ions and molecules are formed to explain how atoms combine. He called his explanation **the octet rule**.

The octet rule states that atoms tend to form compounds in ways that give them eight valence electrons and thus the electron configuration of a noble gas (except hydrogen, helium lithium, beryllium and boron). Hydrogen, lithium and helium become stable when they attain two valence electrons.

Dear learner, in the preceding unit, you have studied the valence electrons of the atoms of elements in each group in a periodic tablen of the element?

The valence electrons of all atoms are between 1 and 8. Most elements follow the octet rule in chemical bonding, which means that an element should have eight valence electrons in a bond or exactly fill up its valence shell. Having eight valence electrons ensures that the atom is stable. This is why noble gases, which have a full valence electron shell, are already stable and tend not to transfer electrons when bonding with another atom.

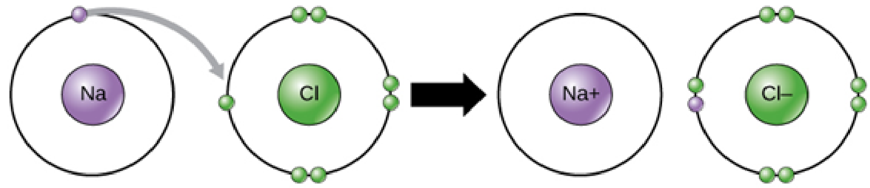
? Dear learner, did you know how atoms could satisfy the octet rule?

There are three ways in which atoms can satisfy the octet rule.

- By losing valence electrons from an atom.
- By gaining valence electrons from other elements.
- By sharing their valence electrons with other atoms.

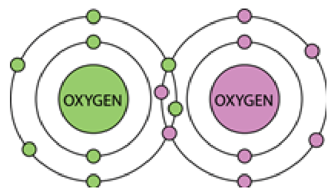
To become stable, atoms of metals tend to lose all of their valence electrons, which leaves them with an octet.

For example, sodium, an alkali metal has 11 electrons and its electronic configuration is 2, 8, and 1. To fulfill the octet rule, sodium needs to lose its outermost electron and hence, its electronic configuration will become 2, and 8.



Atoms of non-metals, on the other hand, tend to gain electrons to fill their outermost principal energy level with an octet.

For example, chlorine is a non-metallic element with 17 electrons. Its electronic configuration is 2, 8, and 7. To fulfill the octet rule, chlorine must gain one electron on its valence shell (see the illustration above). On the other hand, oxygen, being a non-metallic element, has a total of eight electrons. Its electronic configuration will be 2, 6. To attain eight valence electrons, oxygen must gain two additional electrons in its valence shell (see the diagram below). This, therefore, will change its electronic configuration to 2, 8 , which fulfills an octet. Oxygen, however, can do this by sharing its two valence electrons with two electrons of another oxygen atom, which is known as **sharing of electrons**.



Resource

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of why atoms form chemical bonds.

1. https://www.wonderopolis.org/wonder/why-do-atoms-form-molecules

Activity 5.1

1. Define the octet rule.
2. How do metals and nonmetals fulfill the octet rule?
3. How does hydrogen violate the octet rule?
4. Discuss the relationships between an atom, a molecule, a chemical reaction, and a chemical bond.

Self-Test Exercise 5.1

1. What is the electronic configuration of a noble gas?
2. Why is the noble gas configuration important?
3. Do all elements follow the octet rule?
4. What should the elements Mg, Ca, C, and N do to fulfill the octet rule?
5. Why do atoms combine to form molecules of compounds?
6. What should they do to become stable?

Checklist 5.1

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
define chemical bonding?		
describe why atoms form chemical bonds?		

Section 5.2: Ionic Bonding

- Dear learner, in the preceding section, you have studied that for atoms to become stable, they need to gain, lose, or share their valence electrons. This losing, gaining, or sharing of electrons by atoms leads to chemical bonding. The result of chemical bonding will be the formation of molecules or compounds with different chemical and physical properties. In this section, I am going to first discuss the formation of ionic bonding. This will be followed by a discussion about how to write the chemical formulas of ionic compounds using the Lewis formula. The section will also cover the discussion of the properties of ionic compounds.

Learning Outcomes

Dear learner, when you have finished studying this section, you should be able to

- Explain the term ion.
- Elucidate the formation of ions by giving examples.
- Define ionic bonding.
- Describe the formation of an ionic bond.
- Give examples of simple ionic compounds.
- Draw Lewis' structures or electron-dot formulas of simple ionic compounds.
- Describe the general properties of ionic compounds.
- Investigate the properties of given samples of ionic compounds.

Study Notes and Important Points

Dear learner, study the following notes and important points before you do the activity and self-test exercise questions. Make sure that you are clear on the meanings of the concepts, terms, and definitions. After doing so, check whether you achieved the learning outcomes indicated above. This section is further divided into three sub-sections. The first deals with the formation of ionic bonding. The second and third sub-sections are the Lewis formula and general properties of ionic compounds, respectively. The study notes are sequentially presented below.

5.2.1 Formation of Ionic Bonding

Dear learner, let me challenge you with basic questions that will help you understand the formation of ionic bonds. From your lesson on atomic structure in Unit 3, you know that an electrically neutral atom has an equal number of negatively charged electrons and positively charged protons. What would happen to an atom when it gained or lost one electron? What will happen to the charge of atom 'A' if it has one more electron than its protons and atom 'B' if it has one less electron than its protons? Do you think that A and B could be held by a certain force if you put them together? What can you say about the force that holds A and B together?

In Unit Three, you learned that an atom possessing the same number of protons and electrons is electronically neutral. However, if the total number of electrons does not equal the number of protons, the atom has a net positive or negative electrical charge.

? Dear learner, what are ions?

Any atom or molecule with a net charge, either positive or negative, is known as **an ion**.

The positive electric charge of a proton is equal in magnitude to the negative charge of an electron; therefore, the net electric charge of an ion is equal to its number of protons minus its number of electrons.

Net electric charge = number of protons – number of electrons

Ions are highly reactive species. They are generally found in a gaseous state and do not occur in abundance on Earth. Ions in the liquid or solid state are produced when salts interact with their solvents. They are repelled by like electric charges but are attracted to opposite charges.

Dear learner, how many types of ions did you learn in seventh grade science subject?

There are two types of classifications of ions in chemistry. These include:

classification based on the type of electric charge atoms or groups of atoms possess and

classification based on the number of atoms involved in ion formation.

In chemistry, we classify ions based on the type of electric charge of atoms into:

Anions: Atoms or a group of atoms that have more electrons than protons and so have a net negative charge. Examples are Cl^-, I^-, SO_4^{2-}, and NO_3^-.

Cations: Atoms or a group of atoms that have more protons than electrons and so have a net positive charge. Examples are Na^+, Ca^{2+}, and Al^{3+}.

Table 5.1. Cations and anions of some metallic and non-metallic elements.

Element	Symbols	Proton	Electron	Charge	Proton	Electron	Ion
Hydrogen	H	1	1	0	1	0	H^+
Nitrogen	N	7	7	0	7	10	N^{3-}
Oxygen	O	8	8	0	8	10	O^{2-}
Sodium	Na	11	11	0	11	10	Na^+
Magnesium	Ma	12	12	0	12	10	Mg^{2+}
Aluminium	Al	13	13	0	13	10	Al^{3+}
Chlorine	Cl	17	17	0	17	18	Cl^-

? Dear learner, did you know the ways in which ions form?

Ions can be formed by **ionization**, which is the process of a neutral atom **losing or gaining its valence electrons**. Generally, electrons are either added to or lost from the valence shell of an atom. The inner-shell electrons do not participate in this type of chemical interaction. Ionization generally involves the transfer of electrons between atoms or molecules. The process is driven by the achievement of more stable electronic configurations, such as the octet rule. Polyatomic and molecular ions can also be formed, generally by gaining or losing elemental ions, such as H^+, in neutral molecules. Polyatomic ions are generally very unstable and reactive.

Let me briefly discuss ionization using examples. A common example of an elemental ion is Na^+ (read as sodium cation or sodium ion). According to ***Table 5.1***, sodium has a +1 charge when it contains ten electrons (neutral sodium has 11 electrons). However, according to the octet rule, sodium would be more stable with 10 electrons (2 in its innermost shell, and 8 in its outermost shell). We have said that metals lose electrons in order to attain an octet. Therefore, sodium metal tends to lose an electron to become more stable. On the other hand, chlorine, a non-metallic element, tends to gain an electron to become Cl^- (read as a chloride ion). Chlorine naturally has 17 electrons, but it would be more stable with 18 electrons (2 in its innermost shell, 8 in its second shell, and 8 in its valence shell). Therefore, chlorine will take an electron from another atom to become negatively charged. As a general rule, the charge of all group IA metals is 1+, group IIA metals are 2+, group VIIA non-metals are -1. Groups IIIA to VIA will not form an ionic bond in most cases, and they are not the topic of this section.

Ionic Bonding

? Dear learner, so far we have defined ions and chemical bonding. Can you define ionic bonding?

Let us consider the chemical bond that will be formed between sodium metal and chlorine. As we have discussed above, sodium will lose its outermost electron to attain an octet and hence will become a sodium cation (Na^+; 2, 8). Chlorine, being a non-metal, accepts the electron sodium has lost and becomes a chloride ion (Cl^-, 2, 8, 8). When the two oppositely charged ions come together, they will be held together by an electrostatic force of attraction, resulting in the formation of sodium chloride (NaCl), commonly known as table salt.

A bond formed by two oppositely charged ions due to electrostatic force is known as an **ionic bond**.

An ionic bond is also called **an electrovalent bond**. The process of forming such a bond is called **ionic bonding**. ***Figure 5.2*** shows the formation of an ionic bond between sodium and chlorine atoms using atomic diagrams.

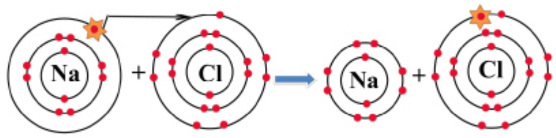
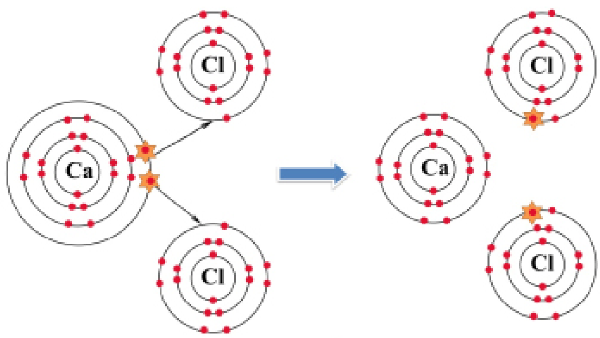


Figure 5.2 An atomic diagram of the formation of an ionic compound (NaCl) between the atoms of sodium and chlorine.

Let us consider a second example, the formation of an ionic bond between calcium and chlorine. Calcium with atomic number 20 has an electronic configuration of 2, 8, 8, 2. To achieve an octet, calcium removes its two valence electrons. This will make calcium a cation (Ca^{2+}; 2, 8, 8) because the number of electrons is less by 2 than the number of protons. Chlorine, as we have discussed above, needs only one electron to fulfil an octet. The two electrons removed from calcium will, therefore, be received by two chlorine atoms to make two chloride ions ($2Cl^-$). One calcium cation (Ca^{2+}) will be attracted by the two chloride ions ($2Cl^-$), resulting in calcium chloride ($CaCl_2$). The atomic diagram in ***Figure 5.3*** depicts the ionic bond formation between calcium and chlorine.



$Ca + 2Cl \rightarrow Ca^{2+} \quad 2Cl^-$

Figure 5.3 Atomic diagram of the formation of the ionic compound $CaCl_2$.

As a general rule, **an ionic bond** is formed between **metallic and non-metallic elements.**

? What is electronegativity?

For atoms with the largest electronegativity differences (such as metals bonding with non-metals), the bonding interaction is called ionic. The valence electrons are typically represented as being transferred from the metal atom to the non-metal. Once the electrons have been transferred to the non-metal, both the metal and the non-metal are considered to be ions. The two oppositely charged ions electrostatically attract each other to form an ionic compound.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of the formation of ionic bonds.

1. https://flexbooks.ck12.org/cbook/cbse-chemistry-class-10/section/3.5/primary/

lesson/ionic-bond-reaction-between-metals-and-nonmetals/
2. https://uen.pressbooks.pub/introductorychemistry/chapter/the-ionic-bond/
3. https://www.youtube.com/watch?v=j5M9_qoGKXY

Activity 5.2

Identify whether the following statements are True or False.

1. Metals gain electrons in order to fulfill the octet rule.
2. Carbon must lose its valence electrons in order to fulfill the octet rule.
3. An element becomes a negatively charged ion if it gains electrons.
4. Chlorine must always share its valence electrons in order to become stable.

Answer the following questions.

5. Define the terms 'cation' and 'anion'.
6. How do cations and anions form?
7. What is a chemical bond?
8. How do ionic bonds form?
9. An aluminum atom has three valence electrons. Do you think it will lose three electrons or gain five electrons to obtain an octet in its outermost electron shell? Why?
10. An iodine atom has seven valence electrons. Do you think it will lose seven electrons or gain one electron to obtain an octet in its outermost electron shell? Why?
11. Describe the formation of potassium iodide (KI).

5.2.2 Lewis Formulas of Ionic Compounds

? Dear learner, can you discuss how an ionic bond is formed? Why do you think it is important to use a common way of representing compounds using formulas?

In this sub-section, you are going to study the Lewis formulas of ionic compounds.
To explain the various types of bonds and to visualise the change in the valence electrons, the American physical chemist Gilbert N. Lewis (1916) proposed the Lewis dot formula.

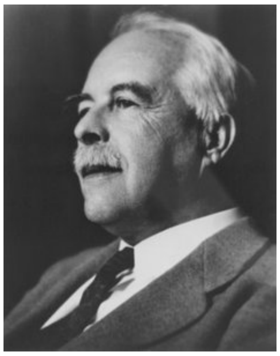


Figure 5.1 The American Physical Chemist Gilbert N. Lewis (1875-1046).

? Dear learner, do you remember what valence electrons are?

In the Lewis symbol for an atom, the chemical symbol of the element with its valence electrons is represented as dots surrounding it. Only the electrons in the valence level are shown using this notation. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right, left, above, and below the symbol, with no more than two dots on a side (*Table 5.2*).

Table 5.2 Lewis symbols for the elements of the third period of the periodic table.

Atom	Electronic configuration	Valence electrons	Lewis's symbol
Sodium	2,8,1	1	$\text{Na}\cdot$
Magnesium	2,8,2	2	$\cdot\text{Mg}\cdot$
Aluminium	2,8,3	3	$\cdot\dot{\text{Al}}\cdot$
Silicon	2,8,4	4	$\cdot\underset{\cdot}{\dot{\text{Si}}}\cdot$
Phosphorus	2,8,5	5	$\cdot\underset{\cdot}{\ddot{\text{P}}}\cdot$
Sulphur	2,8,6	6	$:\underset{\cdot}{\ddot{\text{S}}}\cdot$
Chlorine	2,8,7	7	$:\underset{\cdot\cdot}{\ddot{\text{Cl}}}\cdot$
Argon	2,8,8	8	$:\underset{\cdot\cdot}{\ddot{\text{Ar}}}:$

Lewis symbols can also be used to illustrate the formation of cations from atoms for sodium and calcium, as shown below:

$$\text{Na}\cdot \longrightarrow \text{Na}^+ + e^- \qquad \cdot\text{Ca}\cdot \longrightarrow \text{Ca}^{2+} + 2e^-$$

Likewise, they can be used to show the formation of anions from chlorine and sulphur, as shown below:

$$:\underset{\cdot\cdot}{\ddot{\text{Cl}}}\cdot + e^- \longrightarrow :\underset{\cdot\cdot}{\ddot{\text{Cl}}}:^- \qquad :\underset{\cdot}{\ddot{\text{S}}}\cdot + 2e^- \longrightarrow :\underset{\cdot\cdot}{\ddot{\text{S}}}:^{2-}$$

Table 5.3 Lewis symbols of the transfer of electrons during the formation of ionic compounds and the Lewis formula of some ionic compounds.

Metal		Nonmetal		Ionic Compound
$\text{Na}\cdot$ sodium atom	+	$:\underset{\cdot\cdot}{\ddot{\text{Cl}}}\cdot$ chlorine atom	$\longrightarrow$	$\text{Na}^+\left[:\underset{\cdot\cdot}{\ddot{\text{Cl}}}:\right]^-$ sodium chloride (sodium ion and chloride ion)
$\cdot\text{Mg}\cdot$ magnesium atom	+	$:\underset{\cdot}{\ddot{\text{O}}}\cdot$ oxygen atom	$\longrightarrow$	$\text{Mg}^{2+}\left[:\underset{\cdot\cdot}{\ddot{\text{O}}}:\right]^{2-}$ magnesium oxide (magnesium ion and oxide ion)
$\cdot\text{Ca}\cdot$ calcium atom	+	$2\,:\underset{\cdot\cdot}{\ddot{\text{F}}}\cdot$ fluorine atoms	$\longrightarrow$	$\text{Ca}^{2+}\left[:\underset{\cdot\cdot}{\ddot{\text{F}}}:\right]^-_2$ calcium fluoride (calcium ion and two fluoride ions)

? Dear learner, do you remember how to draw the atomic diagram of elements?

The Lewis dot formula can also be represented using the atomic diagram, as shown in ***Figures 5.2*** (for NaCl) and 5.3 (for $CaCl_2$). The formation of these ionic compounds was discussed in the previous subsection. The dots on the atomic diagram represent the electrons. The starred electron is the electron that is transferred from sodium metal to chlorine in the course of NaCl formation. I have made it a star dot, not because it is different from the other electrons. I just made it a star-dot to show the electron that is transferred from Na and Ca. You don't necessarily need to make a star-dot in order to show the electron that is transferred.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of the Lewis dot formulas of ionic compounds.

1. https://chemistrytalk.org/lewis-dot-structures/
2. https://chem.libretexts.org/Courses/College_of_Marin/CHEM_114%3A Introductory_Chemistry/10%3A_Chemical_Bonding/10.03%3A_Lewis_Structures_of_Ionic_Compounds-_Electrons_Transferred

Activity 5.3

1. Draw the Lewis dot structures for:

   a. MgS , b. Al_2O_3 c. $GaCl_3$
   d. K_2O e. Li_3N f. KF

2. In the Lewis structures listed below, M and X represent various elements in the third period of the periodic table. Write the formula of each compound using the chemical symbols of each element:

   a. $[M^{2+}][:\underset{\cdot\cdot}{\overset{\cdot\cdot}{X}}:]^{2-}$ b. $[M^{3+}][:\underset{\cdot\cdot}{\overset{\cdot\cdot}{X}}:]^{-}_{3}$ c. $[M^{+}]_2[:\underset{\cdot\cdot}{\overset{\cdot\cdot}{X}}:]^{2-}$ d. $[M^{3+}]_2[:\underset{\cdot\cdot}{\overset{\cdot\cdot}{X}}:]^{2-}_{3}$

5.2.3 General Properties of Ionic Compounds

Dear learner, in the previous sub-sections, you have studied the formation of ionic compounds. However, there is one important question that we need to address at this stage, i.e.,

? How can one differentiate ionic compounds from other compounds?

Let me ask you a question that can lead you to the important point of the subsection.

? If you are given juices made up of lemon, grape and orange without letting you know, how do you differentiate the juices?

Now think about how you can differentiate ionic compounds from other compounds.

Ionic compounds contain ionic bonds. An ionic bond is formed when there is a large electronegativity difference between the elements participating in the bond formation. The greater the electronegativity differences among the bonding atoms, the stronger the attraction between cation and anion. The properties of ionic compounds depend on how strongly the cations and anions attract each other in an ionic bond. Based on the above behaviour, ionic compounds exhibit the following properties:

A. They form crystals.

? Dear learner, what do you think is the reason that ionic compounds are crystalline solids?

As you know, substances can exist in one of three states: solid, liquid, or gas. A solid-state substance can exist as a powder, crystal, or amorphous.

? Can you classify corn flour, sugar, and grease into these three forms of solids?

Most ionic compounds exist in crystalline, solid form at room temperature. The constituent particles of the crystals are ions, not molecules.

The oppositely charged ions in ionic compounds are held together very strongly by the electrostatic force of attraction. Hence, ionic compounds are generally hard solids.

At an atomic level, an ionic crystal forms a regular structure, with the cation and anion alternating with each other and forming a three-dimensional structure-based, largely, on the smaller ion uniformly filling in the empty spaces between the larger ions (***Figure 5.4***).

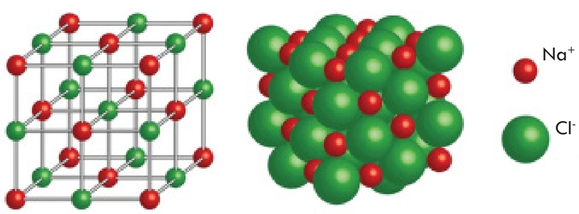


Figure 5.4. Model of the crystal structure of sodium chloride (NaCl).

B. They have high melting and high boiling points.

? Dear learner, do you know what the melting and boiling points of a compound are? What do you think makes ionic compounds high- melting and boiling compounds?

The process of melting an ionic compound requires the addition of large amounts of energy in order to break all of the ionic bonds in the crystal and let the ions move freely. Therefore, a high temperature is required to melt and boil ionic compounds. For example, breaking the ionic bonds in sodium chloride to make it melt requires a temperature of about 800 °C. We call this temperature melting point. It requires a very high –temperature of 1465 °C to boil sodium chloride. Boiling a substance means converting its constituents into a gaseous form.

C. They are hard and brittle

? Dear learner, what does brittleness mean?

Ionic crystals are hard because the positive and negative ions are strongly held together by electrostatic attraction. It takes a large amount of mechanical force, such as striking a crystal with a hammer, to force one layer of ions to shift relative to its neighbour. Such striking brings ions of the same charge next to one another (*Figure 5.5*). The strong repulsive forces between like-charged ions cause the crystal to break.

The property of being hard but liable to be broken is known as **brittleness.**

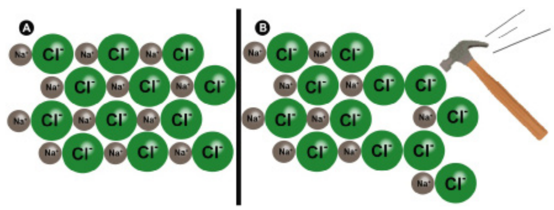


Figure 5.5 Appearance of sodium chloride crystals before and after hitting.

The sodium chloride crystal is shown in two dimensions (*Figure 5.5*, left). When struck by a hammer, the negatively charged chloride ions are forced near one another, and the repulsive force causes the crystal to shatter (*Figure 5.5*, right). Take a crystal of table salt (NaCl) when you go home and hit it with a hummer so that you can see how the crystal breaks.

D. They are soluble in polar solvents

The solubility of a substance is the maximum amount of solute that can dissolve in a given quantity of solvent at a certain temperature and pressure. Solubility depends on the chemical nature of the solute and the solvent, the temperature, and the pressure. One way of classifying solvents is as polar and non-polar. Polar solvents dissolve polar compounds, and non-polar solvents dissolve non-polar compounds. This is commonly referred to as "like dissolves like".

A polar solvent is a type of solvent that has large partial charges.

Water, methanol, liquid ammonia, liquid hydrogen fluoride, acetone, and ethanol are common examples of polar solvents. There are, however, several ionic compounds that are insoluble in water. Most salts of carbonates, oxalates, phosphates, sulphides, sulphates, and hydroxides are insoluble in water. For example, take chalk powder (Ca_2CO_3) and gypsum ($CaSO_4$) and try to dissolve them in water at home.

E. They conduct electricity when they are dissolved in water or in a molten state

? Dear learner, why do you think that ionic compounds do not conduct electricity in the solid state?

Melting and dissolving ionic compounds make cations and anions free to move. Since the ions are free-charged particles, they move towards the respective electrodes under the influence of an electric field and conduct electricity through the solution. A solution of an ionic compound that conducts electricity is known as an electrolyte. For example, sodium chloride (NaCl) in its molten state releases sodium cation (Na^+) and chloride anion (Cl^-), as shown below:

$$NaCl \rightarrow Na^+ + Cl^-$$

Electrolytic cell is a device in which electrical energy is converted to chemical energy, or vice versa.

Such a cell typically consists of positively and negatively charged metallic or electronic conductors (electrodes) held apart from each other and in contact with an electrolyte. In an electrolytic cell (***Figure 5.6***), the sodium cation moves to the negative electrode (cathode) and receives an electron, while the chloride anion moves towards the positive electrode (anode) and gives an electron. This causes the complete movement of charge across the circuit.

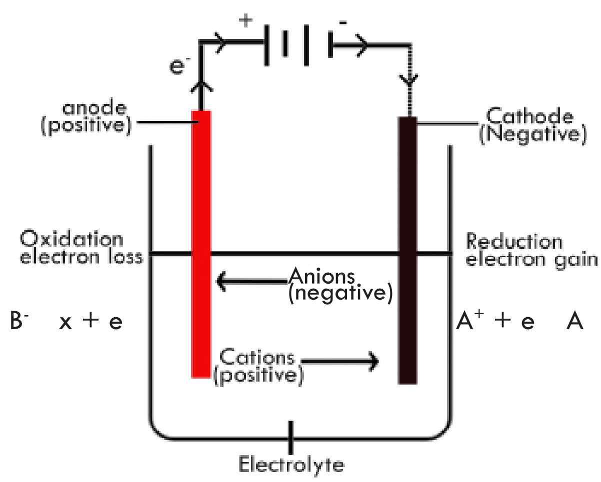


Figure 5.6 An electrolytic cell where the M+ ions migrate toward the negative electrode while the X- ions migrate toward the positive electrode.

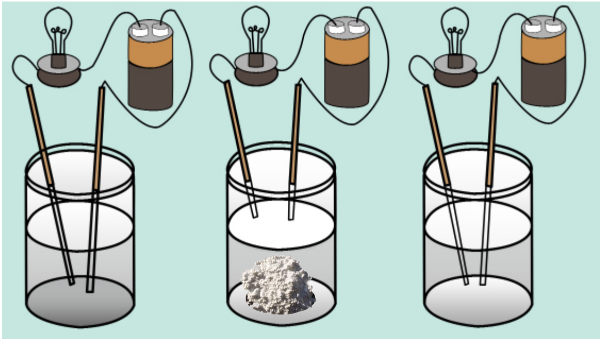
Experiment 5.1: Conductivity of Ionic Compounds

Objective: In this experiment, we are going to see whether distilled water, solid table salt (NaCl), or a water solution of table salt can conduct electricity or not.

Materials: Distilled water (100 mL), table salt (60 g), three 25 mL beakers, battery, copper wire, two pieces of metallic rods or graphite rod, small electric bulb.

Procedure:

1. Take a 25 mL beaker and add 15 mL distilled water.
2. Take 20 g table salt (NaCl) and add it into the second beaker.
3. Dissolve 30 g of table salt (NaCl) in 50 mL distilled water. Add 25 mL of the solution into the third 25 mL beaker.
4. Assemble each of the beakers as shown in ***Figure 5.7*** below, turn by turn and see what happens to the electric bulb in each case.



(a) (b) (c)

Figure 5.7 A) Distilled water; B) Solid table salt powder; C) A water solution of table salt.

Observation and analysis questions

Provide answers to the following questions based on the above experiment.

1. In which case did the electric bulb light?
2. Why the electric bulb did not light in the other two cases?
3. What do you conclude from this observation?

Laboratory report format

Group Number ______________ Date of experiment: ______________________

Name of student(s):

1. ____________________________
2. ____________________________
3. ____________________________

Experiment number: ___________________

Title of experiment: ____

Materials used in the experiment:____

Chemicals: ____

Objective of the experiment: ____

Chemical reaction: ____

Observation:

A. Distilled water: _____________

B. Solid table salt: _____________

C. A water solution of table salt: _____________________

Conclusion: ____

____

____

____.

Answers to observation and analysis questions.

1.

2.

3.

F. They have high density

The density of a substance is the amount of a substance per unit volume. It is one of the physical properties by which substances differ. The oppositely charged ions in an ionic compound are held close by the electrostatic force of attraction. Hence, the number of ions per unit volume in an ionic compound is high, which in turn makes their density high.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of the properties of ionic compounds.

1. https://www.thoughtco.com/ionic-compound-properties-608497
2. https://www.youtube.com/watch?v=0Yj7sRHB5lA
3. https://www.youtube.com/watch?v=SMBA7E6ZZjg

Activity 5.4

1. State the five general properties of ionic compounds.
2. Why do ionic compounds form crystals?
3. If a compound dissolves in water, will it be proof of the ionic nature of the compound? Why?
4. Why do ionic compounds not conduct electricity in the solid state?

Self-Test Exercise 5.2

1. Describe how the ionic compound $MgCl_2$ is formed.
2. Write the Lewis symbols for K, O, and I.
3. Show the formation of K^+, O^{2-}, and I^- using the Lewis formula.
4. Why are ionic compounds brittle?
5. In an electrolytic cell, to which electrode do cations and anions migrate?
6. Why are melting points high for ionic compounds?

Checklist 5.2

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
explain the term ion?		
elucidate the formation of ions by giving examples? define ionic bonding?		
describe the formation of an ionic bond?		
give examples of simple ionic compounds?		
draw Lewis' structures or electron-dot formulas of simple ionic compounds?		
describe the general properties of ionic compounds?		
investigate the properties of given samples of ionic compounds?		

Section 5.3: Covalent Bonding

- Dear learner, in the previous section, you saw that in order to become stable, atoms of elements need to fulfil the octet rule. You have also seen that except for the noble gases, all the second-order raw elements obey the octet rule either by gaining, losing, or sharing electrons. In this section, we shall discuss the second type of bonding known as covalent bonding, the Lewis dot formulas of covalent molecules, polarity in covalent molecules, the coordinate covalent bond, and the general properties of covalent compounds. When you have finished studying this section, you are expected to achieve the following learning outcomes.

Learning Outcomes

Dear learner, when you have finished studying this section, you should be able to

- Define covalent bonding.
- Describe the formation of a covalent bond.
- Draw Lewis' structures or electron-dot formulas of simple covalent molecules.
- Give examples of different types of covalent molecules.
- Make models of covalent molecules showing single, double, and triple bonds using sticks and balls or locally available materials.
- Discuss the polarity of covalent molecules.
- Distinguish between polar and non-polar covalent molecules.
- Define a coordinate covalent (dative) bond.
- Elucidate the formation of coordinate covalent bonds using suitable examples.
- Explain the general properties of covalent compounds.
- Investigate the properties of given samples of covalent compounds.

Study Notes and Important Points

Dear learner, study the following notes and important points before you do the activity and self-test exercise questions. Make sure that you are clear with the meanings of the concepts, terms, and definitions. After doing so check whether you achieved the learning outcomes indicated above. Let me begin this section by discussing the formation of covalent bonds.

5.3.1 Formation of Covalent Bond

? Dear learner, let me begin discussing this subsection by asking you some questions that lead you to the important points of the lesson. What types (metallic and non-metallic) of elements form ionic bonds?. How is it that two non-metallic atoms form a bond? What kind of force holds the atoms together in compounds other than ionic?

A covalent bond consists of the simultaneous attraction of two nuclei for one or more pairs of electrons. In other words, a covalent bond is formed when two atoms share one or more electron pairs. The electrons residing between the two nuclei are known as the bonding electrons. In this type of bond, each shared electron will be counted for both atoms' valence shells in order to satisfy the octet rule. Based on the number of pairs of shared electrons, the covalent bonds formed can be classified into three categories:

Single bond: one pair of electrons is shared between two atoms.
Double bond: the two atoms share two pairs of electrons.
Triple bond: the two atoms share three pairs of electrons.
Generally, bonds sharing more than one pair of electrons are called **multiple covalent bonds.**

Covalent bonds occur between identical atoms or between different atoms whose difference in electronegativity is insufficient to allow the transfer of electrons to form ions.

Formation of a Single Covalent Bond

Let's consider the covalent bond in the hydrogen molecule (H_2). A hydrogen molecule is formed from two hydrogen atoms, each with one electron in its valence shell. The two hydrogen atoms are attracted to the same pair of electrons in the covalent bond. The bond is represented either as a pair of "dots" or as a solid line, between the symbols of the element. Each hydrogen atom acquires a helium-like electron configuration. Such a type of bond is known as a single bond (*Figure 5.8*).

$$H. \quad + \quad .H \quad \rightarrow \quad H{:}H \text{ or } H-H$$

The alternative way of showing the above covalent bond formation would be the following.

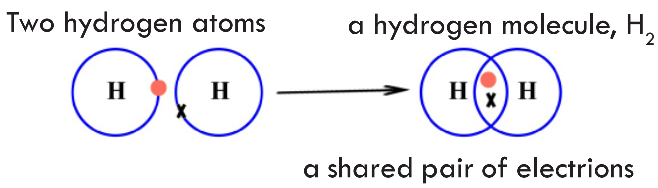


Figure 5.8 Atomic diagram of the formation of a covalent bond in a hydrogen molecule (H_2).

There exists an attraction force between the positively charged nuclei and the negatively charged bonding electrons revolving around the nucleus (*Figure 5.9*). The attractive forces are equal in magnitude but opposite in sign. Each hydrogen nucleus attracts both electrons, and this is the basis of covalent bond formation. The repulsive force between the positively charged nuclei protects the nuclei from collision during covalent bond formation.

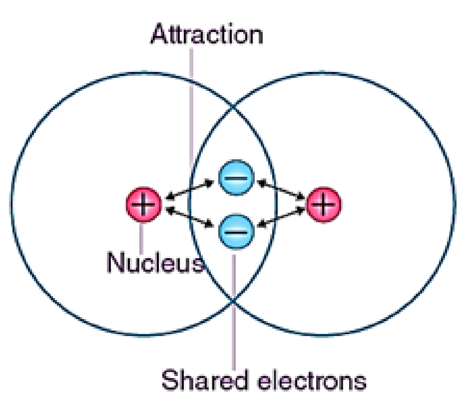


Figure 5.9 Attraction between the nuclei and the electrons in a covalent bond.

Let me briefly describe one more example of single- bond formation. The two chlorine atoms in the chlorine molecule (Cl_2) are joined by a shared pair of electrons. Each chlorine atom contains seven valence electrons in the valance shell and requires one more electron to form an argon-like electron configuration and become stable (***Figure 5.10***). The two chlorine atoms achieve this if each chlorine atom contributes one electron to the bonding pair shared by the two atoms. The remaining six valence electrons of each chlorine atom are not involved in bond formation and are located around their respective atoms. These valence electrons, normally shown as pairs of electrons, are commonly known as nonbonding electrons, lone pair electrons, or unshared electron pairs.

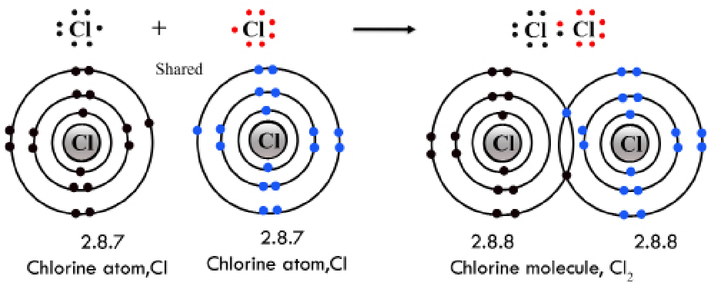


Figure 5.10 Atomic diagram of the formation of a chlorine molecule (Cl_2).

? Dear learner, how about single bond formation between two different atoms?

Single covalent bonds can also occur between two different atoms. For example, hydrogen sulphide (H_2S) is formed when two hydrogen atoms, share their valence shell electrons with one sulphur atom. Sulphur is a group VIA element, and it has six valence electrons. In order to fill its outermost shell with eight electrons, it needs two more electrons, in this case, two valence electrons shared by the two hydrogen atoms. The two hydrogen atoms will have an electronic configuration similar to that of helium, which makes them stable enough, whereas the sulphur atom will attain the electronic configuration of argon, which makes it stable.

? Dear learner, is there another possible way through which two atoms or groups of atoms form a covalent bond other than a single bond?

Let me challenge you with this exercise before directly going into the discussion of this topic. Atom "A" has a total of eight electrons and atom "B" has a total of seven electrons. How do the molecules A_2 and B_2 form? Can you construct a model that shows A_2 and B_2 from locally available materials?

Formation of a Double Covalent Bond

An oxygen molecule (O_2) would be a good example of a double covalent bond formation. Atmospheric air consists of 21% of the life-giving gas known as oxygen. Oxygen is a group VIA element with six valence electrons. Each oxygen atom requires two electrons to fulfil the octet rule. The two oxygen atoms, therefore, need to share two electrons each so that they can attain the nearest noble gas electronic configuration, neon. Since two pairs of electrons are shared between the two oxygen atoms in an oxygen molecule, such a type of bond is known as a double bond. These pairs of electrons can also be represented by two solid lines between the two oxygen atoms, as shown in ***Figure 5.10***. The two pairs of electrons that do not participate in bonding and are situated on the two oxygen atoms in O_2 are the lone pair electrons (***Figure 5.11***).

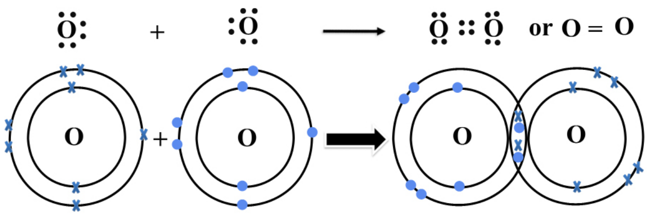


Figure 5.11 Atomic diagram of the formation of oxygen molecules (O_2).

Formation of a Triple Covalent Bond

Nitrogen is a group VA element. It has five valence electrons. We know that 78% of the atmosphere is filled with the element nitrogen. Nitrogen exists in its molecular form, which will not be stable otherwise. This stability could be achieved by fulfilling the octet rule. To do so, the nitrogen atoms share three electrons each to have the nearest noble gas (neon) electronic configuration. The sharing of three pairs of electrons in a covalent bond results in a triple bond. The pairs of electrons on the valence shells of the two nitrogen atoms in N_2 are the lone pair electrons (***Figure 5.12***). A nitrogen molecule could also be represented by N≡N.

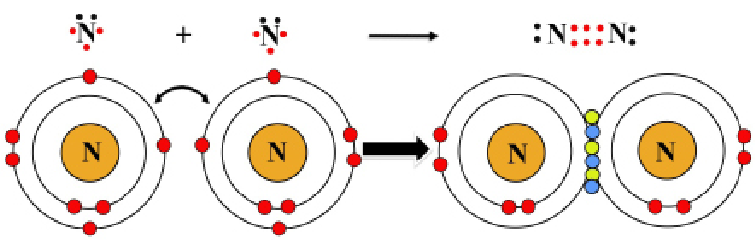


Figure 5.12 Atomic diagram of the formation of nitrogen molecules (N_2).

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of covalent bond formation.

1. https://byjus.com/jee/covalent-bond/
2. https://study.com/learn/lesson/covalent-bonds-examples-formation-properties.html
3. https://www.britannica.com/science/covalent-bond

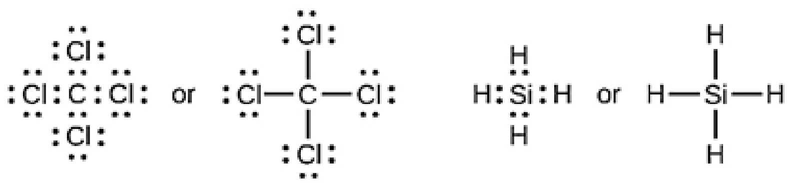
Activity 5.5

1. Define a covalent bond and discuss how it forms.
2. Why do the nuclei of the two covalently bonded atoms not collide during the formation of a covalent bond?
3. In covalently bonded atoms, what holds the two atoms together?
4. Describe the formation of F_2, CF_4, H_2O, and NH_3.
5. How many different types of covalent bonds have you realized so far?
6. Using atomic diagrams, describe the formation of carbon dioxide (CO_2), ethylene (H_4C_2), and acetylene (H_2C_2) molecules.

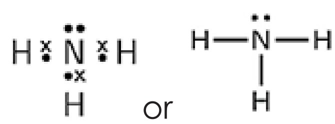
5.3.2 Lewis's Formula of Covalent Molecules

? Dear learner, in the previous subsection, we discussed writing Lewis's formulas for ionic compounds. How can one write Lewis's formula for a covalent molecule?

The number of bonds that an atom can form can often be predicted from the number of electrons needed to reach an octet. This is especially true of the non-metals of the second period of the periodic table (C, N, O, and F). For example, each atom of the group IVA element has four valence electrons and therefore requires four more electrons to reach an octet. These four electrons can be gained by forming four covalent bonds, as illustrated below for carbon in CCl_4 (carbon tetrachloride) and silicon in SiH_4 (silane). Because hydrogen only needs two electrons to fill its valence shell, it is an exception to the octet rule.

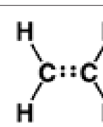
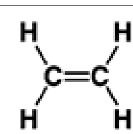


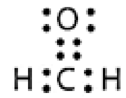
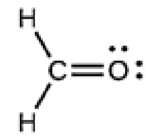
Group VA elements such as nitrogen have five valence electrons in the atomic Lewis symbol: one lone pair and three unpaired electrons. To achieve an octet, these atoms form three covalent bonds, as in NH_3 (ammonia).



Oxygen and other atoms in group VIA achieve an octet by forming two covalent bonds. For very simple molecules, we can write the Lewis structures by merely pairing up the unpaired electrons on the constituent atoms. Look at the examples in ***Table 5.5***.

Table 5.5. Lewis's formula of some simple covalent molecules

Element/molecule	Molecular formula	Dot formula	Line formula
Hydrogen	H_2	H:H	H-H
Nitrogen	N_2	:N:::N:	:N≡N:
Oxygen	O_2	Ö::Ö	:Ö=Ö:
Chlorine	Cl_2	:C̈l:C̈l:	:C̈l—C̈l:
Water	H_2O	H:Ö:H	H-Ö-H
Ethene	C_2H_4		
Acetylene	C_2H_2	H:C:::C:H	H-C≡C-H

Formaldehyde	CH_2O		

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of covalent bond formation.

1. https://byjus.com/jee/covalent-bond/
2. https://study.com/learn/lesson/covalent-bonds-examples-formation-properties.html
3. https://www.britannica.com/science/covalent-bond

Activity 5.6

1. Write the Lewis structure for the diatomic molecule P_2, an unstable form of phosphorus found in high-temperature phosphorus vapour.
2. Write Lewis structures for the following:
   a. O_2
   b. H_2CO
   c. AsF_3
3. Construct models for CH_4, O_2, N_2, and H_2C_2 from locally available materials.

5.3.3 Polarity in Covalent Molecules

? Dear learner, based on your lesson on the periodic table, can you define the term 'electronegativity'? What does the electronegativity difference do to the electron distribution of atoms that are covalently bonded? What would be the electron distribution of a covalent bond formed from two atoms having similar electronegativities? What relationship is there between the polarity of molecules and their properties?

In this sub-section, I will be briefly discussing these points.

The existence of a 100% ionic or covalent bond represents an ideal situation. In reality, no bond or compound is either completely covalent or ionic. Even in the case of a covalent bond between two diatomic molecules, there is some ionic character.

When a covalent bond is formed between two similar atoms, for example in H_2, O_2, Cl_2, N_2, or F_2, the shared pair of electrons is equally attracted by the two atoms (***Figure 5.9***). As a result, the electron pair is situated exactly between the two identical nuclei. The bond so formed is called a nonpolar covalent bond. Contrary to this, in the case of a heteronuclear molecule like HF, the shared electron pair between the two atoms get displaced more towards fluorine since the electronegativity of fluorine (as you have learned in Unit 4) is far greater than that of hydrogen. The resultant covalent bond is a polar covalent bond.

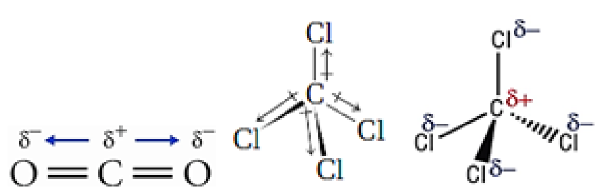
Non-polar covalent molecules

Let us consider the homonuclear diatomic molecule H_2. In a molecule like H_2, in which the atoms are identical, we expect the electrons to be equally distributed between the two atoms. The two hydrogens have equal electronegativity (2.1), and the difference between the electronegativities will be zero. Therefore, the covalent bond in H_2 is non-polar. Generally, all homoatomic molecules like O_2, N_2, and Cl_2 are non-polar covalent molecules.

Let us consider the heteronuclear polyatomic molecule ethane (C_2H_6). There are two different types of covalent bonds in ethane, i.e., H-C and C-C. Considering the electronegativity value difference between hydrogen (2.1) and carbon (2.5),it will be 0.4, which is below the minimum value (0.5) for a polar covalent bond. The C-C bond is non-polar because there is no electronegativity value difference between the two carbons. Hence, ethane is a nonpolar covalent molecule.

Other examples of non-polar molecules include any of the homonuclear diatomic elements (H_2, N_2, O_2, and Cl_2, which are truly non-polar molecules), carbon dioxide (CO_2), benzene (C_6H_6), carbon tetrachloride (CCl_4), methane (CH_4), ethylene (C_2H_4), hydrocarbon liquids (gasoline and toluene), and most organic molecules.

Non-polar molecules also form when atoms share a polar bond arrangement such that the electric charges cancel out each other. For example, in CO_2 and CCl_4, the individual C-O and C-Cl bonds are polar. This is because in the case of the C-O bond, the electronegativity value difference between carbon (2.5) and oxygen (3.5) is 1.0. Hence, the bond is polar. Similarly, in a C-Cl bond, the electronegativity value difference between carbon (2.5) and chlorine (3.0) is 0.5, which is the minimum value requirement for a polar covalent bond. However, in CO_2, the molecular structure is linear, O=C=O and the charges on the two oxygen atoms cancel out each other as they are oriented in opposite directions (see below),making the molecule non-polar. In the case of CCl_4, the chlorine atoms arrange themselves in a tetrahedral geometry (see below) around the carbon atom, and each partial charge on the four chlorine atoms will cancel out each other. This made CCl_4 a nonpolar covalent molecule.



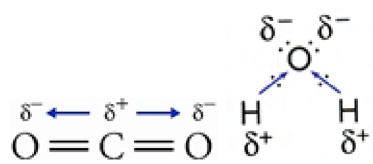
Polar Covalent Molecules

In many covalent bonds that are made up of heteroatoms, the electrons are not shared equally between two bonded atoms. For example, in hydrogen chloride, the electrons are unevenly distributed between the two atoms because the atoms that share the electrons in the molecule are different and have different electronegativities.

Electronegativity is the tendency of an atom to pull bonding electrons toward it.

A bond in which electrons are shared unevenly is called a polar bond or polar covalent bond. A polar bond has a slight positive charge on one end and a slight negative charge on the other end. The greater the difference in electronegativity between the bonded atoms, the more polar the bond will be. The direction of bond polarity can be indicated with an arrow. The head of the arrow is at the negative end of the bond; a short perpendicular line near the tail of the arrow marks the positive end of the bond (see below).

Let us consider one more example: the water molecule H_2O. The electronegativity value difference between oxygen and hydrogen (3.5-2.1 = 1.4) tells us the H-O bond is polar. The magnitude of the electronegativity difference tells us how strongly polar the H-O bond is. Unlike CO_2, the shape of a water molecule is a bent or V-shape. Unlike CO_2, the net charges in H_2O will not cancel each other out. Thus, water is a polar covalent molecule.



Other examples of polar molecules include ammonia (NH_3), sulfur dioxide (SO_2), hydrogen sulfide (H_2S), methanol (CH_4O), and ethanol (C_2H_6O).

The Distinction Between Polar and Non-Polar Covalent Bonds and Ionic Bonds

There is no sharp distinction between a polar covalent bond and an ionic bond, but the following rules are helpful as a rough guide. An ionic bond results when the electronegativity difference between the two bonding atoms is 2.0. This rule applies to most ionic compounds. A polar covalent bond forms when the electronegativity difference between the atoms is in the range 0.5-2.0. If the electronegativity difference is below 0.5, the bond is normally classified as a covalent bond, with little or no polarity. Generally, ionic compounds are highly polar. Atoms of elements with comparable electronegativities tend to form moderately polar covalent bonds with each other because the shift in electron density is usually small. Atoms of the same element, which have the same electronegativity, can be joined by a pure nonpolar covalent bond.

Table 5.6 Relationship between polarity and type of bonding

Electronegativity difference	Bonding	Bond example
0.0 - 0.4	Nonpolar covalent bond	H-C, C-C
0.5 - 0.9	Slightly polar covalent bond	H-N, H-Cl
1.0 - 1.3	Moderately polar covalent bond	C-O, S-O
1.4 - 1.7	Highly polar covalent bond	H-O
1.8 - 2.2	Slightly ionic bond	H-F
2.3 - 3.3	Highly ionic bond	$Na^+ F^-$

Example

Classify the following bonds as ionic, polar covalent, or covalent:

a. The bond in HCl,
b. The bond in KF, and
c. The C-C bond in H_3CCH_3.

Strategy: We follow the 2.0 rule of electronegativity difference and look up the values in the electronegativity value Table.

Solution

a. The electronegativity difference between H and Cl is 0.9, which is appreciable but not large enough (by the 2.0 rule) to qualify HCl as an ionic compound. Therefore, the bond between H and Cl is slightly polar covalent.
b. The electronegativity difference between K and F is 3.2, which is well above the 2.0 mark; therefore, the bond between K and F is highly ionic.
c. The two C atoms are identical in every respect. They are bonded to each other, and each is bonded to three other H atoms. Therefore, the bond between them is purely covalent.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of covalent bond formation.

1. https://www.expii.com/t/polar-vs-nonpolar-bonds-overview-examples-10350
2. https://study.com/learn/lesson/covalent-bonds-examples-formation-properties.html

Activity 5.7

You may refer to the answer key to determine whether your answers are correct or not afterwards. You may need to refer to the internet for the electronegativity values of the elements.

1. Which of the following bonds is nonpolar covalent, which is polar covalent, and which is ionic? (a) the bond in CsCl; (b) the H-N bond in NH_3; (c) the NN bond in H_2NNH_2.
2. Justify your answers to question number 1.
3. Explain why CO_2 is a non-polar whereas SO_2 is a polar covalent bond.

5.3.4 Coordinate Covalent Bond (Dative Bond)

? Dear learner, what are the three ways through which atoms forming chemical bonds could attain the octet rule? Isn't there another possible way for two atoms to form a covalent bond other than the above-mentioned ways? Explain how.

In this section, you are going to study the fourth option of chemical bond formation.
In the previous sections, we have discussed ionic and covalent bonding. There is a fourth possibility in which unstable atoms become stable, and that is by sharing electrons. However, the sharing is only from one atom.

A coordinate covalent bond, or sometimes known as a **dative bond**, is a covalent bond where the electron pair is provided by only one of the bonded atoms but shared by both atoms after bond formation (Figure 5.13).

The atoms are held together because both nuclei attract the electron pair in a similar fashion to that of a covalent bond. Once the coordinate covalent bond is formed, it is impossible to distinguish the origin of the electrons. There are two necessary conditions for a coordinate covalent bond to take place:

One of the atoms in the molecule must have a pair of electrons; in most cases, these are lone pair electrons. Examples of such molecules include H_2O, NH_3, and H_2S.

The other atom in the molecule must have an empty space in its valence shell in order to accept a pair of electrons. Examples of such molecules include BF_3, $AlCl_3$, and H^+.

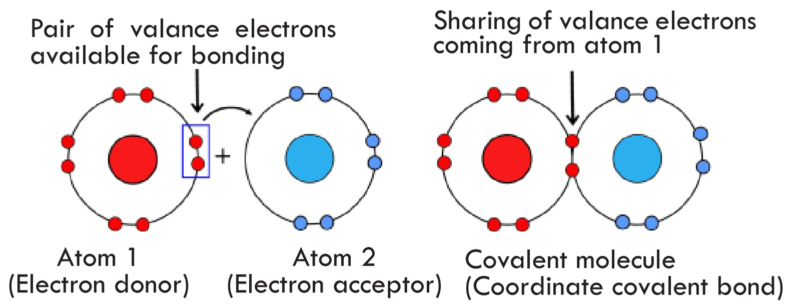
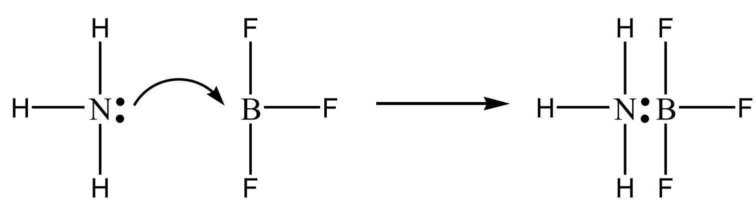
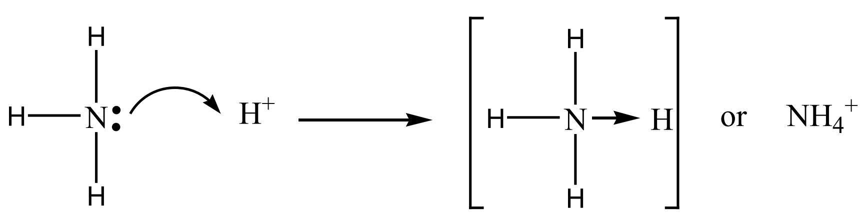


Figure 5.13 Atomic diagram of the formation of a coordinate covalent bond between atoms 1 and 2.

Consider the reaction between BF_3 and NH_3. Boron is one of the elements that become stable without fulfilling the octet rule. In BF_3, the number of valence electrons in boron is only 6. Thus, boron has space to hold two more electrons in its valence shell. On the other hand, in NH_3, nitrogen has a lone pair electron. Although BF_3 is stable, it readily forms a bond with NH_3. This reaction takes place in such a way that nitrogen shares its pair of electrons with boron, forming a coordinate covalent bond as shown below. In this case, both boron and nitrogen share the shared electrons of nitrogen and satisfy the octet rule.



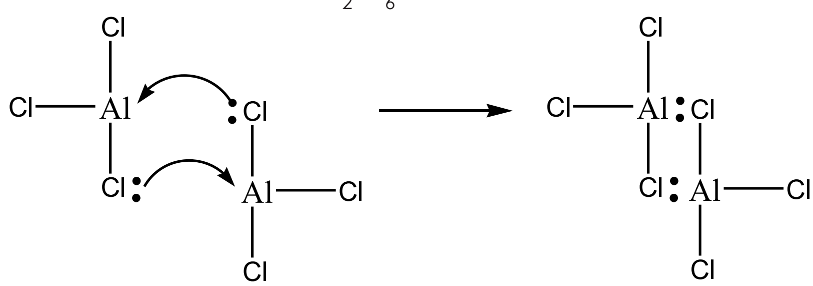
The second example of a coordinate covalent bond is the formation of ammonium ions (NH_4^+) from hydrogen ions (H^+) and ammonia (NH_3). The hydrogen ion has no electrons on its valence shell and can accommodate two electrons in order to attain the electronic configuration of helium and become stable. Ammonia shares its lone pair electrons with hydrogen ions and forms a coordinate covalent bond, as shown below. The coordinate covalent bond is sometimes represented by an arrow pointing from the electron- sharing atom (N in this case) towards the atom having space on the valence electrons (H in this case).



The third example is the formation of Al_2Cl_6. The bonding in aluminum chloride ($AlCl_3$) is essentially covalent.

? Which group element is aluminum in the periodic table of elements? How many valence electrons does it possess?

Each aluminum atom (Al) has a deficit of two electrons in its valance shell in order to fulfill the octet rule, and each chlorine atom (Cl) has a lone pair electron. Aluminum forms a coordinate covalent bond with the Cl atom on an adjacent $AlCl_3$ group, as shown below. As each of two aluminum atoms does this, then aluminum chloride is a covalent dimer molecule with the formula Al_2Cl_6.



Although the properties of a coordinate covalent bond do not differ from those of a normal covalent bond, since all electrons are identical no matter what their source, the distinction is useful for keeping track of valence electrons.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of coordinate covalent bond formation.

1. https://mydigitalkemistry.com/coordinate-covalent-bond-definition-examples-formation-properties/
2. https://www.nagwa.com/en/explainers/235126567926/

Activity 5.8

1. Show the formation of carbon monoxide (CO) using Lewis's formula.
2. Explain the formation of the hydronium ion (H_3O^+).

5.3.5 General Properties of Covalent Compounds

Dear learner, in the previous sub-section, you studied the general properties of ionic compounds. Consider the covalent molecule water (mp 0 °C; bp 100 °C) and the ionic compound table salt (NaCl). Compare them in terms of state, melting point, boiling point, and conductivity of electricity.

? Based on the above comparison, what general properties can you suggest about covalent compounds?

In the previous sub-section, we discussed that a covalent bond is formed by sharing electrons between two atoms. Covalent compounds have different physical and chemical properties compared to ionic compounds. In this section, we are going to discuss the general properties of covalent compounds.

Covalent compounds form discrete molecules that can exist independently from each other. In contrast, ionic compounds have a crystal structures in which formula units cannot exist individually but only as part of a lattice of ions. Therefore, the physical properties of covalent molecules depend heavily on the nature of their interaction with other molecules (intermolecular forces). Depending on the nature of these intermolecular interactions, covalent compounds have the following properties.

At room temperature, most covalent compounds are gases or liquids. Some covalent compounds are soft solids.

The intermolecular force between most covalent molecules is the weak van der Waal's force and will be discussed in Grade 11.

Most covalent compounds have low melting and boiling points.

The melting point and boiling point of substances are also dependent on the intermolecular forces between covalent molecules. Diamond is an exception. Its melting point is very high (about 4027 °C).

Most covalent compounds are poor conductors of electricity.

This is because they cannot form ions in their solution form or molten state. Graphite, an allotrope of carbon, has covalent bonds and is an exception as it is a good conductor. Some covalent compounds, like HCl that ionise in an aqueous solution are good conductors of electricity as well.

Most covalent compounds are soluble in non-polar solvents and are insoluble in polar solvents like water.

The reason behind this is most covalent compounds are either non-polar or moderately polar. Water is a highly polar solvent and cannot dissolve non-polar or moderately polar covalent molecules. Few polar covalent compounds, however, dissolve in polar solvents such as methanol and ethanol. The principle "**like dissolves like**" meaning polar solvents dissolve polar compounds and non-polar solvents dissolve non-polar compounds, works in the solubility of substances.

Reactions of covalent compounds are slow compared to those of ionic compounds. This is because of the existence of bond breaking in chemical reactions, where breaking a covalent bond requires high energy. In reactions involving ionic compounds, the dissolution of ionic compounds releases ions relatively easily. Hence, a relatively fast reaction occurs.

The existence of a covalent compound in the liquid or gaseous state makes the number of molecules per unit volume less, thereby leading to low density.

Intermolecular forces and their relationship to the states of matter will be discussed in Grade 11.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of the general properties of covalent compounds.

1. https://www.thoughtco.com/covalent-or-molecular-compound-properties-608495
2. https://www.toppr.com/guides/chemistry/chemical-bonding-and-molecular-structure/covalent-compounds/

Activity 5.9

1. Describe the physical properties of covalent compounds.
2. Compare the physical properties of covalent compounds with those of the ionic compounds.

Self-Test Exercise 5.3

Choose the correct answer from the given alternatives.

1. Which of the following compounds contains both covalent and ionic bonds?
   A. NaOH C. NaNC
   B. NaBr D. NaCN
2. A chemical bond between two atoms that share a single pair of electrons is
   A. ionic bond C. double Bond
   B. single bond D. triple Bond
3. Which of the following compounds contains both polar and non-polar covalent bonds?
   A. NH_4Br C. CH_4
   B. H_2O_2 D. HF

Answer the following questions.

4. Describe the electronegativity difference between each pair of atoms and the resulting polarity (or bond type).
   A. C and H C. Na and Cl
   B. H and H D. O and H
5. 5.Determine which atom in each pair has the higher electronegativity.
   A. H or C C. Na or Rb
   B. O or Br D. I or Cl
6. Will the electrons be shared equally or unequally across each covalent bond? If unequally, to which atom are the electrons more strongly drawn?
   A. C–O bond C. S–N bond
   B. F–F bond D. I–Cl bond
7. Arrange the following bonds from least polar to most polar: C-N, C-O, C-C, C-H, N-H, O-H
8. Explain how nitric acid (HNO_3) is formed.

Checklist 5.3

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark. Can I ...

Competencies	Yes	No
define covalent bonding?		
describe the formation of a covalent bond?		
draw Lewis' structures or electron-dot formulas of simple covalent molecules?		
give examples of different types of covalent molecules?		
make models of covalent molecules showing single, double, and triple bonds using sticks and balls or locally available materials?		
discuss the polarity of covalent molecules?		
distinguish between polar and non-polar covalent molecules?		
define a coordinate covalent (dative) bond?		
elucidate the formation of coordinate covalent bonds using suitable examples?		
explain the general properties of covalent compounds?		
investigate the properties of given samples of covalent compounds?		

Section 5.4: Metallic Bonding

Overview

? Dear learner, what types of bonding have you studied so far? From which types (metallic, non-metallic) of elements each type of bond is formed? How does each type of bonding occur?

These, however, are not the only types of bonding atoms that can form. There is a third type of bonding that occurs between metallic atoms. Unlike other types of bonds, this type of bond will only occur between atoms of the same element. This section, therefore, deals with the formation of metallic bonding and its properties.

Learning Outcomes

Dear learner, when you have finished studying this section, you should be able to

- Discuss the formation of a metallic bond.
- Explain the electrical and thermal conductivity of metals in relation to metallic bonding.
- Make a model to demonstrate metallic bonding.

Study Notes and Important Points

Dear learner, study the following notes and important points before you do the activity and self-test exercise questions. Make sure that you are clear on the meanings of the concepts, terms, and definitions. After doing so, check whether you achieved the learning outcomes indicated above. This section will be further divided into two sub-sections known as formation of metallic bond, and properties of metallic bond. Let me begin with the formation of a metallic bond.

5.4.1 Formation of Metallic Bond

? Dear learner, why do you think there are no metal molecules? How do metallic elements exist in their pure form? From your lesson on the periodic table, how many valence electrons are there in metallic elements in general?

A metal atom generally has 1, 2, or 3 electrons in its valence shell. It can easily lose these electrons, and gain stability in the course of a chemical reaction. These electrons lost by the metal are called **free electrons**. Metals are thus highly electropositive in nature. **Electropositivity** is the tendency to lose electrons and form positive ions in chemical reactions. Metallic bonds occur among the same metal atoms. A sheet of aluminium foil and a copper wire are good examples where you can see metallic bonding in action.

Sodium metal has the electronic configuration 2, 8, 1. In sodium metal, the electron in the outermost shell of one sodium atom shares space with the equivalent valence electron on a neighbouring atom's outermost shell. This happens in much the same way that a covalent bond is formed. The difference between a covalent bond and the metallic bond in sodium atoms, however, is that each sodium atom is touched by eight other sodium atoms, and the sharing occurs between the central atom and the outermost shell on all of the eight other atoms. Each of these eight atoms, in turn, is being touched by eight sodium atoms, which in turn are touched by eight atoms, and so on and so on, until all the atoms in that lump of sodium are taken. All of the outermost shells on all of the sodium atoms overlap to give a vast number of molecular shells that extend over the lamp of metal.

The electrons freely move within these molecular outermost shells, and therefore each electron becomes detached from its parent atom. The electrons are said to be **delocalized**. The metal is held together by the strong forces of attraction between the positive nuclei and the delocalized electrons (***Figure 5.14***). Such a type of electrostatic attraction is called a **metallic bond**. In this figure, the orange circles with the "+" sign at the centre represent the positive atomic nuclei of the metal, and the small dotted yellow circles represent a sea of delocalized electrons.

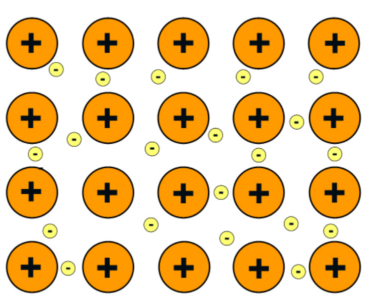


Figure 5.14 The Electron Sea metallic bonding model

Such type of metallic bonding is sometimes described as "an array of positive ions in a sea of electrons".

Therefore, **metallic bonding** is a type of bonding that has an array of the same positive metal ions in a sea of delocalized electrons.

Each positive centre in the diagram (***Figure 5.14***) represents all the rest of the atom apart from the outer electron, but that electron has not been totally lost from the atoms. It may no longer have an attachment to a particular atom, but it's still there in the structure. Sodium metal is therefore written as Na, not Na^+.

Problem: Discuss the metallic bonding in magnesium and explain why it has a higher melting point than sodium metal.

Strategy: Use the sea of electrons model to explain why magnesium has a higher melting point (650 °C) than sodium (97.79 °C).

Solution

Magnesium has the electronic configuration 2, 8, 2. Both of the valence electrons become delocalized, so the "sea" has twice the electron density as it does in sodium. The remaining "ions" also have twice the charge and so there will be a stronger attraction between "ions" and "sea". Each magnesium atom has 12 protons in the nucleus, compared with sodium's 11. In both magnesium and sodium, the nucleus is screened from the delocalized electrons by the same number (10) of electrons in the inner shells (2 + 8). This means that there will be a stronger attraction from the magnesium nucleus of “2+” than from the sodium nucleus, which has only a “1+” nucleus.

So, there will be a greater number of delocalized electrons in magnesium, which leads to greater attraction by the magnesium nuclei. Furthermore, the smaller radius of magnesium atoms leads to stronger attraction of the delocalized electrons than sodium, as they are closer to the magnesium nuclei. Each magnesium atom also has a greater number (twelve) of near neighbours than sodium (eight). These factors, therefore, increase the strength of the bond even further in magnesium metal.

Transition metals tend to have particularly high melting points and boiling points. The reason is that they can involve more valence electrons in the delocalization. The more electrons you can involve, the stronger the attraction tends to be and the stronger the metallic bond will be.

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of metallic bonding.

1. https://www.expii.com/t/metallic-bond-formation-compounds-8645
2. https://www.youtube.com/watch?v=-ntOkgo7knY
3. https://www.youtube.com/watch?v=Bjf9gMDP47s

Activity 5.10

1. Define the metallic bond and explain how it forms.
2. Discuss the formation of metallic bonds in aluminium metal. (Hint: One aluminium atom will be surrounded by 12 atoms).
3. Compare the metallic strength between magnesium and aluminium metals. Justify your answer.
4. Draw the electron sea model of magnesium metal.

5.4.2 Properties of Metallic Bond

Dear learner, in the previous sub-section, you studied the way metallic bonding occurs in metallic elements. Let me begin this section by asking a question that leads you to the main point of the subsection.

? Which of the following substances conduct electricity: dry cloth, plastic, dry wood, or a piece of metal? Why? Which substance is the hardest? Why?

Metals have several unique qualities, such as the ability to conduct electricity and heat. Metals have other bulk properties, which will be discussed in Grade 10.

Conductivity: Since the electrons in metals are free to move between the positively charged "ions", pushing electrons from an outside source (e.g., battery) into a metal wire at one end (-Q; ***Figure 5.15***) would move the free electrons (the small bluish circles) through the wire and come out at the other end (+Q) at the same rate. Such a movement of charge is known as conductivity. These freely moving electrons transfer electric charge and heat across the metallic structure. This freedom of the valence electrons accounts for the high thermal conductivity of metals as well.

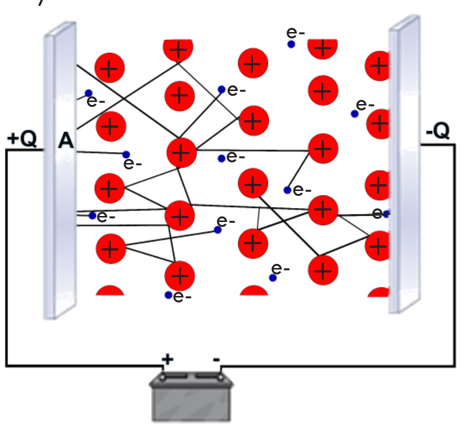


Figure 5.15 The "sea of electrons" (small bluish circles) is free to flow about the crystal of positive metal ions (the large reddish circles).

Resources

Dear learner, please go to the internet and copy and paste these URLs (website links) so that you can get a better understanding of the properties of metallic bonds.

1. https://www.expii.com/t/metallic-bond-formation-compounds-8645
2. https://www.youtube.com/watch?v=-ntOkgo7knY
3. https://www.youtube.com/watch?v=Bjf9gMDP47s

Activity 5.11

1. Why do metals conduct heat and electricity?
2. What are the three factors that determine the strength of a metallic bond?
3. Why do metals have high melting and boiling points compared to non-metals?

Self-Test Exercise 5.4

Choose the correct answer for the following questions.

1. Metallic bonding describes the bonds between two metals.
   A. True B. False
2. Metals have high melting and boiling points.
   A. True B. False
3. In metals, the ___________ electrons form a shared sea of electrons.
   A. metallic
   B. inner
   C. outer
   D. ionic
4. In general, what can be said of the melting point of K?
   A. It is low
   B. It is lower than that of Ca
   C. It is greater than the value of Mg
   D. It is the same as Na
5. The electron sea metallic bonding model is . . . . ?
   A. The model of metallic bonding is one where electrons float free in a sea of electrons around metal atoms.
   B. The model of metallic bonding is one where electrons are fixed in place in a sea of metal atoms.
   C. A model depicting the different bonds that electrons can make.
   D. The model of metallic bonding is one where protons float free in a sea of electrons around metal atoms.
6. What does the 'sea of electrons' contain?
   A. All the electrons in that metal.
   B. The electrons in the outer shell of that metal.
   C. The electrons in the inner shell of that metal.
   D. The electrons in the penultimate shell
7. Why are metals able to conduct electricity?
   A. The positive metal ions pass charges between each other.
   B. Electrons pass charges through the positive metal ions.
   C. The sea of electrons helps pass charges through the metal.
   D. The positive nuclei of the metals move through the electron sea.

Answer the following questions.

8. In a metallic bond, what held the metal atoms together?
9. Compare the metallic bond strength of K with that of Ca.

Checklist 5.4

Dear learner, check yourself on the following aspects before wrapping up studying this section. In the table given below, put "✓" in the "Yes" column if you are quite sure that you have done it and "X" in the "No' column if you are not sure or did not do it. You must answer all the questions and revise those with an "X' mark.

Competencies	Yes	No
discuss the formation of a metallic bond?		
explain the electrical and thermal conductivity of metals in relation to metallic bonding?		
make a model to demonstrate metallic bonding?		

Unit Summary

In this unit, the association of the electrons and the properties of the elements in chemical bond formation are discussed thoroughly. Elements are distinguishable from each other due to their electrons. Because each element has a distinct number of electrons, this determines its chemical properties as well as the extent of its reactivity. In chemical bonding, only valence electrons are involved. This unit, therefore, dealt with the way the different types of chemical bonds are formed, the properties of compounds formed in each type of chemical bonding, the electron dot formulas or Lewis structures used to represent simple molecules or compounds formed in each type of chemical bonding, and the origin of polarity in molecules.

Except for a few elements, the atoms of most of the elements have a characteristic tendency to combine and form molecules or compounds. The atoms that form the molecules are held together by attractive forces. **A chemical bond** is, therefore, the force that holds atoms together to form molecules or compounds.

The octet rule states that atoms tend to form compounds in ways that give them eight valence electrons, thus the electron configuration of a noble gas. Most elements follow the octet rule in chemical bonding, which means that in order to become stable, an element should have eight valence electrons in a bond or exactly fill up its valence shell (except H, He, and group IIIA elements).

In the course of chemical bond formation, atoms attain an octet by sharing, losing, or gaining valence electrons. Metals tend to lose all their valence electrons, while nonmetals gain or share some of their valence electrons.

The net electric charge of an ion is calculated as follows:

Net electric charge = number of protons – number of electrons

Anions are atoms or groups of atoms having a net negative electric charge and cations are atoms or groups of atoms having a net positive electric charge. Ions can be formed by ionization, which is the process of a neutral atom losing or gaining its valence electrons.

The electrostatic attraction of two oppositely charged ions creates what is called **ionic bond**. The process of forming such a bond is called **ionic bonding**. As a general rule, an ionic bond is formed between metallic and non-metallic elements.

In the Lewis symbol for an atom, the chemical symbol of the element with its valence electrons is represented as dots surrounding it. The Lewis dot formula can also be represented using the atomic diagram.

The properties of ionic compounds depend on how strongly the cations and anions attract each other in an ionic bond. Ionic compounds form crystals, have high melting and boiling points, are hard and brittle, are soluble in polar solvents, conduct electricity when dissolved in water or in a molten state, and have a high density.

A covalent bond is formed when two atoms share one or more electron pairs. Based on the number of pairs of shared electrons, the covalent bonds formed can be classified into three categories:

Single bond: one pair of electrons is shared
Double bond: share two pairs of electrons and
Triple bond: share three pairs of electrons

In a heteronuclear molecule like HF, the shared electron pair between the two atoms gets displaced more towards fluorine since the electronegativity of fluorine is far greater than that of hydrogen. The resultant covalent bond is a **polar covalent bond**. The molecule is termed as a **polar molecule**.

An ionic bond results when the electronegativity difference between the two bonding atoms is 2.0 or more. A polar covalent bond forms when the electronegativity difference between the atoms is in the range 0.5-2.0.

A **coordinate covalent bond** sometimes known as a dative bond is a covalent bond where the electron pair is provided by only one of the bonded atoms but shared by both atoms after bond formation. Once the coordinate covalent bond is formed, it is impossible to distinguish the origin of the electrons.

Depending on the nature of intermolecular interactions, covalent compounds are gases or liquids, have low melting and boiling points, are poor conductors of electricity, are soluble in non-polar solvents and insoluble in polar solvents, are less reactive than ionic compounds, and have a low density.

Metallic bonds occur among the same metal atoms. In metallic bonding, the metal is held together by the strong forces of attraction between the positive core and the delocalized valence electrons of the metal. Metals have several unique qualities, such as the ability to conduct electricity and heat.

Self-Assessment Exercise

Dear learner, there are about 53 questions categorized into basic, intermediate, and challenge levels. The questions will help you check your understanding of the unit as a whole. You are expected to give answers to all questions before referring to the answers given at the end of the unit. Follow the instructions provided for each level of question and respond accordingly.

Part I: Basic- level questions

Choose the correct answer for the following questions.

1. Which of the following does not contain an ionic bond?
   A. Sulfur dioxide
   B. Sodium oxide
   C. Silicon dioxide
   D. Silver oxide
2. Which substances have ions in their bonding models?
   A. Copper
   B. Carbon dioxide
   C. Copper oxide
   D. Carbon
3. Which of the following is an ionic compound?
   A. HCl
   B. NaCl
   C. CO_2
   D. HBr
4. Which of the following doesn't obey the octet rule?
   A. CH_4
   B. CCl_4
   C. BCl_3
   D. CO_2
5. Which of the following compounds does not contain lone-pair electrons?
   A. H_2O
   B. NH_3
   C. HF
   D. CCl_4

Part II: Intermediary level questions.

Tell whether the following ttatements are True or False.

6. An Octet rule states that an element should have eight valence electrons in a bond or exactly fill up its valence shell.
7. Many elements in the third-row and beyond have been observed to form compounds in which the central atom is surrounded by more than eight electrons.
8. An ionic compound is denoted by writing its net negative charge in superscript immediately after the chemical structure of the atom or molecule.
9. Ionization, which is the process of an atom losing or gaining electrons, can result in the formation of ions. As a general rule, an ionic bond is formed between non-metallic elements.
10. An ionic bond forms when the electronegativity difference between the two bonding atoms is less than 2.0.
11. The properties of covalent compounds relate to how strongly the positive and negative ions attract each other in a covalent bond.
12. Most of the ionic compounds are crystalline solids.
13. Considerable heat energy is required to separate the strong electrostatic attraction force between the positive and negative ions in ionic compounds.
14. The number of bonds that an atom can form can often be predicted from the number of electrons needed to reach an octet.
15. A non-polar covalent bond forms when the electronegativity difference between the atoms is in the range of 0.5-2.0.
16. Once the covalent bond is formed, it is possible to distinguish the origin of the electrons.

Fill in the blank spaces.

17. There are two ways in which atoms can satisfy the octet rule. One way is by ______________ their valence electrons with other atoms. The second way is by ______________ valence electrons from one atom to another.
18. Losing, gaining or sharing of the valence electrons is always accompanied by____________.
19. Atoms of ___________ tend to lose all of their valence electrons, which leaves them with an octet from the next lowest principal energy level.
20. Any atom or molecule with a net charge, either positive or negative, is known as _____________.
21. Two oppositely charged ions due to an electrostatic force create a bond known as_____________.
22. In the _____________ for an atom, the chemical symbol of the element is written, and the valence electrons are represented as dots surrounding it.
23. Polar solvents dissolve polar compounds, and non-polar solvents dissolve non-polar compounds. This is commonly referred to as ___________.
24. Bonds sharing more than one pair of electrons are called ______________.
25. _____________ occur between identical atoms or between different atoms whose difference in electronegativity is insufficient to allow the transfer of electrons to form ions.
26. Only atoms of the same element that have the same electronegativity can be joined by a _____________ bond.
27. A property that helps us distinguish a non-polar covalent bond from a polar covalent bond is _________________.
28. When molecules share electrons equally in a covalent bond, there is no net electrical charge across the molecule. Such a type of covalent molecule is known as a____________ molecule.
29. A _______________ is a type of covalent bond where the electron pair is provided by only one of the bonded atoms but shared by both atoms after bond formation.

Part III: Challenge-Level Questions

Give Appropriate Answers to the Following Questions.

30. How many types of chemical bonds are there? What are they?
31. Name a solvent in which most of the ionic compounds dissolve.
32. In which state do ionic compounds conduct electricity?
33. Why do non-polar covalent compounds not conduct electricity?
34. What type of chemical bond is found in each of the following compounds?
    a) Potassium chloride
    b) Carbon dioxide
    c) Hydrogen chloride
    d) Water
    e) Magnesium oxide
    f) Calcium fluoride
    g) Methane
    h) Sodium chloride
    i) Ammonia
    j) Phosphorus pentachloride
    k) Sulphur hexachloride

35. In what types of solvents do a) polar compounds and b) nonpolar compounds dissolve?
36. Why are molecules more stable than atoms?
37. What are the criteria by which a covalent bond becomes polar or non-polar?
38. Why are partial positive and negative charges developed within a polar covalent molecule?
39. What is a coordinate covalent bond?
40. Define a) polar covalent compounds and b) non-polar covalent compounds.
41. Give some examples in which coordinate covalent bond formation takes place.
42. Draw the Lewis dot formulas for the bond formations in the following compounds.

A.	NaCl	D.	O_2	G.	NH_3
B.	CaF_2	E.	N_2	H.	CH_4
C.	H_2	F.	H_2O		

43. What are the factors responsible for the formation of the covalent bond and the ionic bond?
44. Why do noble gases not take part in a chemical reaction?
45. What type of bond formation takes place between a) a metal and a non-metal and b) two non-metals?
46. Why do most of the ionic compounds exist in the solid state while the covalent compounds exist mostly in a gaseous or liquid state?
47. Why is the density of ionic compounds high and that of covalent compounds low?
48. Why are the melting and boiling points of ionic compounds high and those of covalent compounds low?
49. Why do pure covalent compounds not conduct electricity?
50. Why are ionic compounds soluble in water?
51. Why do polar covalent compounds dissolve in water?
52. What types of bonds exist in the following ions? a) ammonium ion b) hydronium ion?

Assignment for Submission

You are required to give answers to all the questions given below and submit them to your tutor or marker. There are eight 'fill in the blank' items, 17 'multiple choice', and six 'short answer' type questions. All questions are related to the contents of this unit. Do not forget to keep a copy for yourself.

Part I. Fill in the blank spaces.

Fill in the blank spaces of the following questions.

1. Answer the following questions based on the formation of sodium oxide.

   The electronic configuration of the sodium atom is __________________.

   Sodium atom loses its valance electron to possess a______________ charge.

   The electronic configuration of sodium ions is ____________________.

   The electronic configuration of the oxygen atom is _____________________.

   The oxygen atom gains ____________________ electrons from the sodium atom and possesses ________________ charge.

   The electronic configuration of an oxide ion is ________________________.

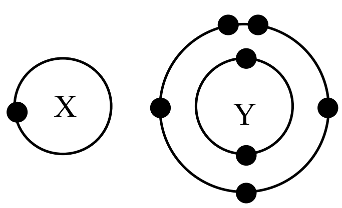
The formula of the ionic compound sodium oxide is ____________________.

2. _______________ is an atom that becomes electrically charged by gaining or losing an electron.
3. When an atom ____________ an electron, it loses a negative charge and becomes a ___________ ion.
4. When an atom ____________ an electron, it gains a negative charge and becomes a _____________ ion.
5. ______________ tend to lose electrons in chemical reactions, forming positive ions.
6. ______________ tend to gain electrons in chemical reactions, forming negative ions.
7. Gaining or losing neutrons makes an atom ________________. However, gaining or losing a _____________ makes an atom into a completely different element.
8. A _________ covalent bond results when electrons are shared equally. A _________ covalent bond results when electrons are shared unequally.

Part II. Multiple Choices

Choose the correct answer for each question.

9. Metal P reacts with non-metal Q to form a compound. Which process takes place and what type of bond is formed?
   A. Electrons are transferred from P to Q, and a covalent bond is formed.
   B. Electrons are transferred from P to Q, and an ionic bond is formed.
   C. Electrons are transferred from Q to P, and a covalent bond is formed.
   D. Electrons are transferred from Q to P, and an ionic bond is formed.
10. The electronic structures of atoms X and Y are shown below. If X and Y form a covalent compound, what is the formula of the compound formed?



   A. XY_5 C. XY
   B. XY_3 D. X_3Y
11. Magnesium atoms react with an oxygen molecule. What is the formula of the product?
   A. MgO C. Mg_2O
   B. MgO_2 D. Mg_2O_2
12. The table below shows the number of valence electrons in four atoms. Which two atoms combine to form a covalent bond?

Atom	Valence electrons
W	1
X	4
Y	7
Z	8

   A. W and X C. X and Y
   B. W and Y D. X and Z

13. Element X is in group IA of the periodic table. X reacts with element Y to form an ionic compound. Which equation shows the process that takes place when X forms ions?

A. $X + e- \rightarrow X^-$
B. $X - e- \rightarrow X^+$
C. $X \rightarrow X^+ - e-$
D. None

14. Sodium chloride is an ionic solid. Which statement is not correct?

A. Ions are formed when atoms loose or gain electrons.
B. Ions in sodium chloride are strongly held together.
C. Ions with the same charge attract each other.
D. A sodium chloride solution can conduct electricity.

15. Rubidium is in group IA of the periodic table, and bromine is in group VIIA. Rubidium reacts with bromine to form an ionic compound. Which one of the following shows the electron change taking place for bromine and the correct formula of the bromide ion?

A. Electrons are gained and Br+ is formed
B. Electrons are gained and Br- is formed.
C. Electrons are lost and Br+ is formed
D. Electrons are lost and Br- is formed.

16. From the combination of atoms below, which one cannot form an ionic bond?

A. K and F
B. Mg and Cl
C. H and Br
D. Ca and I

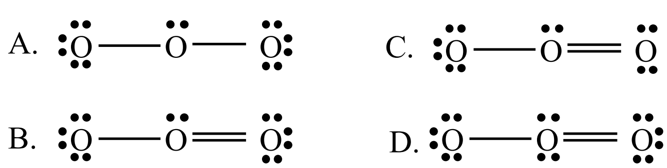
17. For which substance is the type of bonding not correct?

Choice	Substance	Type of bonding		
		Ionic	covalent	metallic
A	Chlorine molecule		✓	
B	Potassium bromide	✓		
C	Sodium			✓
D	Calcium chloride		✓	

18. The most plausible Lewis structure for a chlorate ion, ClO_3^-, should contain _______ single bond(s), ___________ double bond(s), and ___________ lone pair(s) of electrons.

A. 2, 1, 10
B. 1, 2, 8
C. 3, 0, 10
D. 2, 1, 9

19. Which Lewis structure best represents the ozone molecule, O_3?



20. The formal charge on the bromine atom in BrO3- drawn with three single bonds is

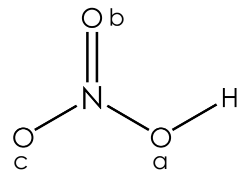
A. -2
B. -1
C. +1
D. +2

21. Nitrous oxide, or N_2O, is sometimes called "laughing gas". What is the formal charge on the oxygen atom in the most plausible Lewis structure for nitrous oxide (the atom connectivity is N-N-O)?

A. -1
B. 0

C. +1 D. +2

22. Nitric acid, HNO_3, is massively used in the production of explosives and fertilizers. Based on the Lewis structure of nitric acid, state the formal charge of each element in the compound as shown in the table below.



Choice	H	N	O_a	O_b	O_c
A	0	-1	0	0	+1
B	-1	0	+1	+1	-1
C	0	+1	0	0	-1
D	+1	+1	-1	0	-1

23. From the following species, which one does not form a dative bond?
A. H_3O^+
B. HO_2^-
C. NH_4^+
D. Al_2Cl_6

24. The electronic configurations of elements X and Y are as follows: X = 2, 3; Y = 2, 8, 7. When X and Y combine, the most plausible formula and bonding are
A. XY_3 – covalent
B. X_2Y – covalent
C. XY_2 – ionic
D. X_2Y_3 – ionic

25. Which element can form an incomplete octet?
A. boron
B. bromine
C. bismuth
D. barium

Part III. Essay

Write your answer to each question below.

26. Write three differences between ionic and covalent bonds.

Covalent	Ionic

27. Fill in the table below.

Aluminum		Oxygen	
Electronic configuration		Electronic configuration	
Valence electrons		Valence electrons	
Aluminum ions formation reaction		Oxide ion formation reaction	
Electronic configuration of aluminum ion		Electronic configuration of oxide ions	
Lewis formula for aluminum ion		Lewis formula for oxide ions	
Formula of aluminum oxide			

28. The table below shows the physical properties of compounds X and Y. One of them is ionic, and the other is covalent. Answer the questions based on the data in the table.

Compound	Melting point
X	2470
Y	801

A. State the type of chemical bond in compounds X and Y.
B. State the species formed in compounds X and Y.
C. Explain why compound X has higher melting point than compound Y.

29. What determines the polarity of a chemical bond?
30. Note that every molecule with polar bonds is polar. Explain this using CCl_4.
31. Explain why metals have high melting and boiling points.

Answer Key to Exercises

Answers to Activity 5.1

1. Octet rule: Atoms tend to form compounds in ways that give them eight valence electrons and thus the electron configuration of a noble gas.
2. Metals lose their valence electrons to fulfil the octet rule. Non-metals fulfil the octet rule by gaining electrons. Sometimes they obey the octet rule by sharing electrons.
3. Hydrogen can fill its outermost shell by gaining or sharing one electron, which will make the number of valence electrons two. Hydrogen will become stable with this number of electrons rather than eight, according to the octet rule. This is how hydrogen violates the octet rule.
4. An atom is the smallest particle of matter that takes part in a chemical reaction. A molecule is the smallest constituent of a substance that has an independent existence, and which represents the properties of the respective elements or compounds. Molecules are made up of atoms through a chemical reaction. A chemical bond is a force that holds atoms together to form molecules or compounds.

Answers to Self-Test Exercise 5.1

1. The electronic configuration of a noble gas is the configuration wherein the number of electrons in the valence shell is eight or two.
2. The noble gas configuration is important to determine the stability of an atom.
3. Only most of the second- raw elements follow the octet rule.
4. In order to fulfil the octet rule, Mg and Ca should lose two of their outermost shell electrons; C should either gain or lose four electrons; and N should gain three electrons.
5. Atoms combine in order to form molecules or compounds because they are not stable or they do not fulfill the octet rule.
6. In order to be stable, atoms must have the electronic configuration of noble gas elements. To do this, they must gain, lose, or share their valence electrons.

Answers to Activity 5.2

True-False

1. False
2. False
3. True

4. False

Short answer

5. A cation is a positively charged atom or group of atoms. An anion is an atom or a group of atoms with a negative charge.
6. Cations are formed when an atom loses its outermost electrons. Anions are formed when an atom gains an electron in its outermost shell.
7. A chemical bond is a force that binds atoms or a group of atoms together.
8. The attraction between oppositely charged ions forms the ionic bonds.
9. It is relatively easier for aluminium to lose three valence electrons than to gain five electrons in its valence shell. This is because it takes more ionization energy for aluminium to gain five electrons than to lose three electrons.
10. It is relatively easier for iodine to add one electron to its valence shell than to lose its seven valence electrons, which require large ionization energy.
11. Potassium loses its valence electron and becomes K^+. Iodine accepts the electron released by potassium and becomes I-. In doing so, both potassium and iodine will attain an octet and become stable. The electrostatic attraction between the potassium cation and iodide ion will form the ionic bond in potassium iodide (KI).

Answers to Activity 5.3

1. Lewis' symbols

Ionic compound	Lewis' symbol (formula)
a. MgS	$Mg^{2+}\ :\ddot{\underset{\cdot\cdot}{S}}:^{2-}$
b. Al_2O_3	$2Al^{3+}\ 3\ :\ddot{\underset{\cdot\cdot}{O}}:^{2-}$
c. $GaCl_3$	$Ga^{3+}\ 3\ :\ddot{\underset{\cdot\cdot}{Cl}}:^{-}$
d. K_2O	$2K^{+}\ :\ddot{\underset{\cdot\cdot}{O}}:^{2-}$
e. Li_3N	$3Li^{+}\ :\ddot{\underset{\cdot\cdot}{N}}:^{3-}$
f. KF	$K^{+}\ :\ddot{\underset{\cdot\cdot}{F}}:^{-}$

2. A group of eight elements, namely, sodium (Na), magnesium (Mg), aluminium (Al), silicon (Si), phosphorous (P), sulphur (S), chlorine (Cl), and argon (Ar) that belong to the third period of the periodic table are called third-row elements. The formula:

A. $MgCl_2$ C. Na_2S

B. $AlCl_3$ D. Al_2S_3

Answers to Activity 5.4

1. The general properties of ionic compounds are

   They form crystals.

   They're hard and brittle.

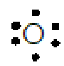
   They have high melting and boiling points.

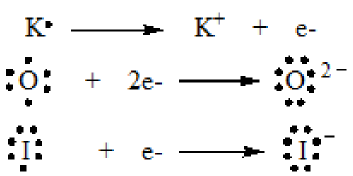
   They conduct electricity when they are dissolved in water and in the molten state.

   They have high density.

2. Ionic compounds are formed from cationic and anionic species that are strongly held by electrostatic force, which makes the ions have a regular, packed structure known as a crystal.
3. No. Because some non-ionic compounds are also soluble in water. For example, sugar.
4. This is because when they are in the molten state or an aqueous solution, the ions become free to move, and hence, they carry an electric charge. This is not true when they are in the solid state, i.e., when the ions are strongly held together in the crystal lattice and cannot move freely.

Answers to Self-Test Exercise 5.2

1. Magnesium has 12 electrons. Its electronic configuration is 2, 8, 2. It is metal and therefore gives its two valence electrons to the two chlorine atoms, which have 17 electrons, and electronic configuration of 2, 8, and 7. This makes the electronic configuration of the chlorines 2, 8, and 8. Magnesium becomes Mg^{2+}, and the chlorines become $2Cl^-$. This, therefore, results in $MgCl_2$.
2. The Lewis symbols: Potassium K•; Oxygen ; Iodine 
3. The fomrmation of K^+, O^{2-}, and I^-



4. The negatively charged ions in the crystal will be forced near one another, and the repulsive force will cause the crystal to break.
5. The cations move towards the negatively charged electrode (cathode) and accept electron(s), whereas the anions migrate to the positively charged electrode (anode) and discharge electron(s).
6. Cations and ions that are held together by strong electrostatic force forms ionic compounds. Breaking this strong force requires a high temperature.

Answers to Activity 5.5

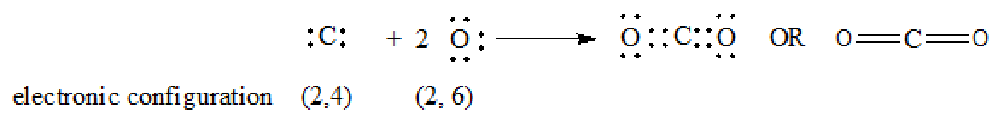
1. A covalent bond is a type of chemical bond formed by the sharing of electrons between two identical or different atoms.
2. The nuclei of the two atoms are positively charged and therefore repel each other. This repulsion protects the nuclei from collision.
3. This is because of two reasons: i) the valence electrons that are found on the two covalently bonded atoms possess a negative charge; hence, they repel each other, and ii) the nuclei of the two atoms possess a positive charge ;hence, there is electrostatic repulsion between them. There is an optimum nuclear distance for the two atoms to exist as a molecule.
4. Formation of F_2: Fluorine, being a group VIIA, element has seven valence electrons. In order to fulfil the octet rule, both electrons must gain one electron. This will be possible if both atoms share one of their valence electrons.

Formation of CF_4: Carbon is a group IVA element, and it has four valence electrons. It needs four extra electrons to fulfil the octet rule. Fluorine is a group VIIA element, and has seven valence electrons, but it needs one extra electron to fulfil the octet rule. Carbon and fluorine can form a covalent bond if carbon shares its four valence electrons with four fluorine atoms, and each fluorine atom shares one of its valence electrons with carbon.
Formation of H_2O: Hydrogen is a group IA element and has one valence electron. It needs one extra electron in order to have the electronic configuration of helium, which makes it stable. Oxygen, being a group ViA element, has six valence electrons and needs two extra electrons to fulfil the octet rule. Hydrogen and oxygen can form a covalent bond when the two hydrogen atoms share their electrons with oxygen, and oxygen also shares two of its valence electrons with the two hydrogens.

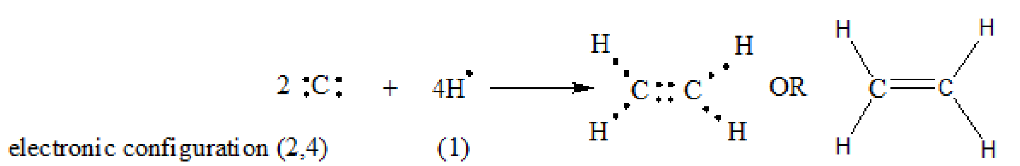
Formation of NH_3: Nitrogen is a group VA element and has five valence electrons. It needs three extra electrons in order to fulfil the octet rule. Nitrogen therefore shares three of its valence electrons with the three hydrogen atoms, and the hydrogens also share their electron with nitrogen.

5. There are three types of covalent bonds: single bonds, double bonds, and triple bonds.
6. Formation of:

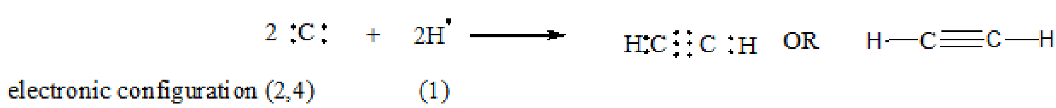
CO_2 =



H_4C_2 =

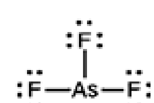


H_2C_2 =



Answers to Activity 5.6

1. P_2: $:P\equiv P:$
2. The Lewis structures are as follows:

a. O_2: $:\ddot{O}=\ddot{O}:$ b. H_2CO:  c. AsF_3: 

3. Make sure that you made the models of the given molecules from locally available materials.

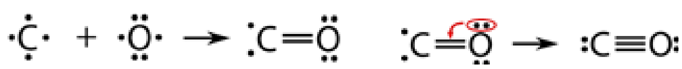
Answers to Activity 5.7

4. Polar covalent bond: NH_3; ionic bond: CsCl; non-polar covalent bond: N-N bond in H_2NNH_2.

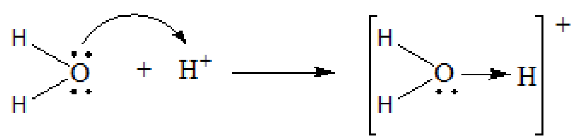
5. The H-N bond in NH_3 is polar because the electronegativity difference between H (2.1), and N (3.0) is 0.9 and since this is greater than 0.5, it qualifies to be polar. The N-N bond is a non-polar covalent bond because there is no electronegativity difference between the nitrogen atoms. CsCl is an ionic compound because i) the electronegativity difference between Cs (0.7) and Cl (3.0) is 2.7 and qualifies to be an ionic bond. ii) Cs is a metal and Cl is a non-metal; hence, the bond between metallic and non-metallic elements is ionic.
6. Although both C-O and S-O bonds are polar, in CO_2, the two oxygen atoms arrange themselves linearly, which cancels out each other's partial charges. In SO_2, however, the oxygen atoms arrange themselves in a bent shape, and the partial charges will not cancel out each other.

Answers to Activity 5.8

1. Carbon (C) has four electrons in its valence shell, and oxygen (O) has six. Both carbon and oxygen share two electrons. While the Octet rule is satisfied with oxygen, there is still a deficit of two electrons on carbon. So, oxygen shares its two electrons with carbon to form a coordinate covalent bond, in addition to the two regular (double) covalent bonds.



2. When hydrogen chloride (HCl) gas dissolves in water to make hydrochloric acid (HCl aq.), a coordinate covalent bond is formed in the hydronium ion. The hydrogen (H) nucleus is transferred to the water (H_2O) molecule, which has a lone pair of electrons, to form hydronium. So, H does not contribute any electrons to the bond.



H_2O molecule H^+ ion Hydronium ion

Answers to Activity 5.9

1. The physical properties of covalent compounds:

   Most covalent compounds are gases or liquids at room temperature. Some are soft solids.

   Most covalent compounds have low melting and boiling points.

   Most covalent compounds are poor conductors of electricity.

   Most covalent compounds are soluble in non-polar solvents and insoluble in polar solvents like water.

   Reactions of covalent compounds are slow compared to those of ionic compounds.

   They have low density.

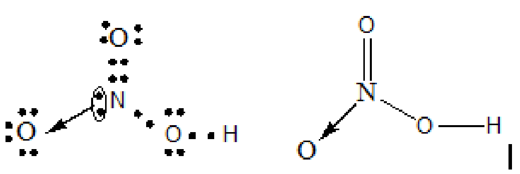
2. Comparison of ionic and covalent compounds:

Property	Ionic compound	Polar covalent compound	Nonpolar covalent compound
Physical state	Crystalline solids. The constituent particles of the crystals are ions, not molecules.	Generally, liquids or gases. The constituent particles are molecules.	
Melting and boiling points	Have high melting and boiling points as a result of the need for considerable heat energy required to overcome the electrostatic force between ions.	Have low melting and boiling points because a small amount of energy is sufficient to overcome the weak electrostatic force of attraction and hydrogen bonding between the polar molecules.	Have low melting and boiling points because a small amount of energy is sufficient to overcome the weak intermolecular forces acting between the molecules.
Solubility	Generally soluble in polar solvents like water, but insoluble in nonpolar organic solvents.	Soluble in polar solvents due to the presence of partial charges. Also soluble in nonpolar covalent liquids, due to similar forces between the molecules.	Insoluble in polar solvents like water because they don't ionize, but soluble in nonpolar covalent liquids like benzene, CCl_4, due to similar forces.
Density	The oppositely charged ions in an ionic compound are held close by an electrostatic force of attraction. Hence, the number of ions per unit volume in an ionic compound is higher, and thereby their density is high.	Generally, they exist in the form of liquid or gaseous states due to weak intermolecular forces. Hence, the number of molecules per unit volume is less, thereby leading to low density.	
Electrical conductivity	Electrovalent compounds conduct electricity either in the fused state or in their aqueous solutions due to the presence of molten ions.	These compounds ionize in water and the ions help conduct electricity.	These compounds ionize in water and the ions help conduct electricity. These compounds do not ionize; hence, do not conduct electricity.

Answers to Self-Test Exercise 5.3

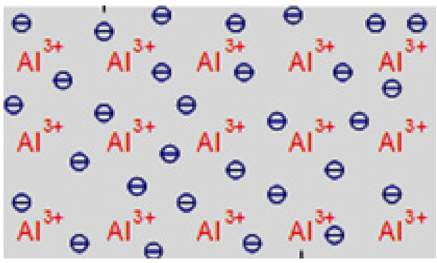
1. Answer c: A covalent bond is present between the N and C atoms and an ionic bond is present between the Na+ ion and the –NC ion.
2. Answer b: Single bond
3. Answer b: In H_2O_2, the electronegativity difference between the O and H atoms is 1.4, and O–H bond is polar. The electronegativity difference between the O and O bonds is zero, so the O – O bond is non-polar.

4. The electronegativity difference between each pair of atoms and the resulting polarity (or bond type):
   a. Carbon has an electronegativity of 2.5, while the value for hydrogen is 2.1. The difference is 0.3, which is rather small. The C–H bond is therefore considered nonpolar.
   b. Both hydrogen atoms have the same electronegativity value—2.1. The difference is zero, so the bond is nonpolar.
   c. Sodium's electronegativity is 0.9, while chlorine's is 3.0. The difference is 2.1, which is rather high, and so sodium and chlorine form an ionic compound.
   d. With 2.1 for hydrogen and 3.5 for oxygen, the electronegativity difference is 1.4. We would expect a very polar bond, but not so polar that the O–H bond is considered ionic.
5. The atoms with the higher electronegativity in each pair are a) C; b) O; c) Na; and d) Cl.
6. The electrons will be shared: a. unequally toward the O; b. equally; c. unequally toward the N; and d. unequally toward the Cl.
7. The electronegativity difference increases from 0; 0.4; 0.5; 0.9; 1.0; 1.4. Hence, the order from least to most polar will be C-C, C-H, C-N, N-H, C-O, O-H.
8. The nitrogen atom makes an ordinary single covalent bond with one oxygen atom and it makes a double covalent bond with a second oxygen atom. The nitrogen atom also makes one coordinate covalent bond as it donates a lone pair of electrons to a third oxygen atom.



Answers to Activity 5.10

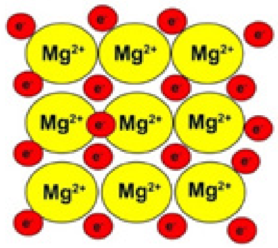
1. The force of attraction that exists between the delocalized electrons and the metal ions is known as a metallic bond.
2. The electronic configuration of aluminium is 2, 8, 3. The three valence electrons in every 12 atoms of aluminium will be freed to move among the pool of 13 protons containing a positively charged aluminium nucleus. The attraction force between the positively charged aluminium nuclei and the shared free electrons makes the metallic bonding in aluminium.



3. The size of Al^{3+} is smaller than that of Mg^{2+}. There are more free electrons in aluminium

than in magnesium. The free electrons in aluminium are, therefore, closer to the positively charged nucleus, thereby forming a strong electrostatic attraction. On the other hand, the size of Mg^{2+} is relatively large, and the number of free electrons is smaller compared to that of aluminium. The attraction force between the freely moving electrons and the nucleus in magnesium is not as strong as that between aluminium's nucleus and the electrons.

4. The electron sea model of magnesium metal. The dots in the red circle represent electrons.



Answers to Activity 5.11

1. Metals conduct heat and electricity because they have free electrons that could easily move in between the nuclei of the metals. Since electrons can move freely, they can easily transfer energy and electricity through the metal.
2. The strength of a metallic bond depends on three things:
   - The number of electrons that become delocalized from the metal.
   - The charge of the cation (metal).
   - The size of the cation.
3. Metallic bonds are strong and require a great deal of energy to break, so metals have high melting and boiling points.

Answers to Self-Test Exercise 5.4

1.	a	4.	b	7.	C
2.	a	5.	a		
3.	c	6.	B		

8. The positively charged nuclei of the metal atoms attract the electron sea delocalized around the positively charged nuclei. It is this electrostatic attraction force that holds the metal atoms together.
9. Potassium (K) is a group IA metal and has one valence electron that creates a sea of electrons delocalizing around the positively (1+) charged potassium nuclei. Calcium (Ca), on the other hand, is a group IIA metal with two valence electrons that create a sea of electrons delocalizing around the positively charged (2+) calcium nuclei. According to Coulomb's law, the attraction force between two charges depends on the amount of charge. Since the charge on calcium metal and its delocalized electrons is greater than the charge on potassium metal and its delocalized electron, the metallic bond in calcium is stronger than that of potassium.

Answers to Self-Assessment Questions

Part I: Answers to the basic- level questions.

1.	a	3.	b	5.	d
2.	c	4.	c		

Part II: Answers to the intermediate-level questions.

True-False questions

6. True
7. True
8. False
9. True
10. False
11. False
12. False
13. True
14. True
15. True
16. False
17. False

Fill in the blank.

18. Sharing, transferring
19. energy changes
20. Metals
21. anion
22. ionic bond
23. Lewis symbol
24. "like dissolves like"
25. multiple covalent bonds
26. Covalent bonds
27. pure covalent
28. electronegativity
29. non-polar
30. coordinate covalent bond or dative bond

Part III. Answers to the challenge level questions.

31. There are four types of chemical Bonds. These are ionic bonds, covalent bonds, coordinate covalent bonds, and metallic bonds.
32. The solvent in which most of the ionic compounds dissolve is water.
33. Ionic compounds conduct electricity in the molten state and in solution form.
34. The non-polar covalent compounds do not conduct electricity because the net charge formed during chemical bond formation is zero. In the absence of a charge carrier, electric conduction is impossible.
35. The type of chemical bond found in each of the following compounds is
    a. Potassium chloride - ionic
    b. Carbon dioxide - covalent
    c. Hydrogen chloride - covalent
    d. Water - covalent
    e. Magnesium oxide - ionic
    f. Calcium fluoride - ionic
    g. Methane - covalent
    h. Sodium chloride - ionic
    i. Ammonia - covalent
    j. Phosphorus pentachloride – covalent
    k. Sulphur hexachloride - covalent
36. According to the principle, "Like dissolves like," a) The polar compounds dissolve in polar solvents b) The non-polar compounds dissolve in non-polar solvents.
37. Molecules are more stable than atoms because the atoms that make molecules attain octate. In other words, the valence shell of the atoms in molecules is filled either with two or eight electrons, which makes them stable. All atoms except for the noble gas elements have a valence electron short of 8 and are not stable.
38. A covalent bond becomes non-polar whenever the atoms forming the bond are identical, meaning homoatomic, or when the electronegativity difference between the atoms is low (less than 0.5). The covalent compound becomes polar when the electronegativity difference between the combining atoms is large (between 0.5-2.0).
39. In a polar covalent molecule, partial positive and negative charges develop because the shared electrons distribute unevenly between the bonding atoms. This is because

of the electronegativity difference between the atoms. In other words, the one with high electronegativity will attract the shared electrons towards itself more than the one with low electronegativity, resulting in an uneven electron distribution.

40. A coordinate covalent bond is a type of bond formed by the sharing of electron pairs between two atoms, but the sharing comes from one atom.
41. A polar covalent compound is a compound in which there is a net partial positive and negative charge in the covalent bond of the molecule. A non-polar covalent compound is a compound in which there is no net partial positive or negative charge in the overall covalent bond of the molecule.
42. Examples of molecules or ions in which coordinate covalent bond formation takes place are the following: H_3O^+, NH_4^+, NH_3BF_3, Al_2Cl_6, and CO.
43. The Lewis' dot formula of the given molecules/compounds is given in the table below:

Element/molecule	Dot formula
NaCl	$Na^+[:\ddot{\underset{\cdot\cdot}{Cl}}:]^-$
CaF_2	$Ca^{2+}[:\ddot{\underset{\cdot\cdot}{F}}:]_2^-$
H_2	H:H
O_2	$\ddot{\underset{\cdot\cdot}{O}}::\ddot{\underset{\cdot\cdot}{O}}$
N_2	:N:::N:
H_2O	$H:\ddot{\underset{\cdot\cdot}{O}}:H$
NH_3	$H:\overset{H}{\underset{\cdot\cdot}{\ddot{N}}}:H$
CH_4	$H:\overset{H}{\underset{H}{C}}:H$

The factors responsible for the formation of covalent and ionic bonds are ionization potential and electron affinity.

44. Noble gas elements do not take part in a chemical reaction because they fulfil the Octet and Duet rules. This means they have a stable valence shell electronic configuration, and hence, they are sufficiently stable.
45. A type of bond formation that takes place between a) a metal and a non-metal is an ionic bond and b) two non-metals is a covalent bond.
46. The reason that most ionic compounds exist in a solid state is because of the strong electrostatic force that exists between them. Covalent compounds have a weak electrostatic or van der Waal's force between the molecules; hence, they are liquids or gases in most cases.
47. The oppositely charged ions in an ionic compound are held close by the electrostatic force of attraction. Hence, the number of ions per unit volume in an ionic compound is higher, making their density high. In covalent molecules, the attraction force between molecules is either a hydrogen bond, or a dipole-dipole interaction, and is not comparable to that of the electrostatic attraction between ions. Hence, most covalent compounds are liquids or gases and it is known that liquids and gases have

low densities.

48. Why are the melting and boiling points of ionic compounds high and those of covalent compounds low? The strong electrostatic force between ionic compounds requires high energy. This high heat energy is associated with high temperature. Hence, ionic compounds need a high temperature to melt or boil. In covalent compounds, the relatively weak dipole-dipole, the hydrogen bonding, or van der Waal's force, needs relatively less energy, which in turn needs a low temperature to melt or boil.
49. Pure covalent compounds have zero net charges. In the absence of a charge, electricity cannot be conducted. In other words, pure covalent compounds do not ionize.
50. Both ionic compounds and water are very polar, and since like dissolves like, polar compounds dissolve in a polar solvent. Ionic compounds exist in the form of ions and dissolve in the highly polar water molecule.
51. Polar covalent compounds dissolve in water because they have a partial positive and negative charge that can interact with the partial negative and positive charges in the water molecules.
52. Types of bonds exist in a) ammonium ion – the covalent bond between three hydrogens and nitrogen and a coordinate covalent bond between the fourth hydrogen and nitrogen; b) hydronium ion – covalent bond between two hydrogens and oxygen and a coordinate covalent bond between the third hydrogen and oxygen.