CHEMISTRY

FOR

Senior Secondary School

1

EDUBASE

S.S.S 1

CHEMISTRY

FIRST TERM

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Week 1

Topic: Introduction to Chemistry

Meaning of Chemistry

Chemistry is a branch of science that deals with the composition, properties and the uses of matter. It probes into the principles governing the changes that matter undergoes. Our environment is made up of matter; chemistry therefore deals with the study of our environment and explaining things that are happening in the environment.

Chemistry may be divided into three parts;

- Inorganic Chemistry: This deals with the study of matter in our environment which is non-living.
- Organic Chemistry: This deals with the study of matter found in living things both plants and animals.
- Physical Chemistry: This deals with the study of energy changes accompanying transformation of matter.

Studying Chemistry

Our world is made up of matter. We study chemistry to acquire knowledge about matter. We perform experiments and learn how to observe, record and make intelligent inferences. Studying chemistry gives us training in the scientific method. This knowledge and training will help us to become scientists in the field of chemistry.

Chemistry in Our Life

Chemical changes happen around us all the time. Examples are

1. Lighting a match
2. Cooking
3. Burning fire wood
4. Making palm wine
5. Rusting nails

Careers in Chemistry

Many job opportunities are available for students with knowledge of chemistry. The public and private sectors of the economy which offer such opportunities are:

Ø Teaching Service

Ø Health Service

Ø Food Processing

Ø Petroleum and Petrochemical industries

Ø Manufacturing Industries

Ø Agriculture and Forestry

Uses of Chemistry

Chemistry has contributed greatly towards providing our basic needs and improving the quality of our life.

1. Food: Fertilizers and insecticides have helped to increase food production greatly. It has also helped in the preservation and storage of food for long periods.
2. Clothing: A wide range of clothing materials are made available because of the man-made fibres produced as a result of intensive research.
3. Military and Space Science: Chemistry contributes to the discovery of chemical substances such as explosives used by the military. It has also helped in an effort to gain more knowledge of the other planets and outer space around us.
4. Housing: The mechanical properties of building materials such as cement, concrete, steel, bricks and tiles are a result of chemical research.
5. Medicine: The healthy life that many of us enjoy is due to the variety of medicines that are available as a result of chemical research and technology.
6. Transportation: This rapid development in transportation, from carts pulled by animals to the latest aircraft was made possible by chemists producing suitable fuels and structural materials.

Adverse effect of Chemistry

1. Corrosion of Metals – This can also be called rusting and requires the presence of water. It can be regarded as slow deterioration of Iron and this can be observed on our Iron sheets at the top of our buildings after exposure to constant rain
2. Pollution – This is caused by chemical industries who through their various processe emit dangerous chemicals into the environment of man. e.g Oil spillage, release of Carbon mono-oxide, pesticides, fertilizers

3. Drug Abuse – Drugs like heroin, cocaine and morphine are addictive. Although these drugs are not used in Medical treatment, unscrupulous people produce and sell them at huge amounts for profit

Assessment

1. Chemistry is defined as
   a. a branch of knowledge which produces chemicals
   b. a branch of science which makes physics and biology clearer
   c. the oldest branch of science
   d. the branch of science which deals with changes in matter

2. Chemical changes around us include all except .....
   a. rusting of iron nails
   b. fading of coloured cloth
   c. sieving
   d. decomposing of green leaves in a compost

3. One of these professions as no need for chemistry
   a. miners
   b. engineers
   c. philosophers
   d. geologists

4. One of these is not a chemical change
   a. rusting
   b. sublimation of solids
   c. slaking of quick lime
   d. fermentation of glucose

Answers

1. D

2. B

3. C

4. D

Week 2

Topic: Nature of Matter

What is Matter?

Matter is anything that has mass and occupies space. There are 3 states of matter, solids, liquids and gases. Examples of matter include the plants and animals around us, the food we eat, the water we drink and even the air we breathe. In general, matter is built up of one or more of the following elementary particles: atoms, molecules and ions.

Properties of Matter

Substances can be identified using the characteristics possessed by the substance. These characteristics are called properties. e.g sugar is a white solid that dissolves in water and tastes sweet. The properties may be physical or chemical.

- Physical properties are properties associated with physical changes. Common physical changes of a substance include its boiling point, melting point, density, volume, mass, hardness, malleability, crystalline form, as well as properties which may be detected by the senses such as colour, odour, and taste.
  All properties of matter are either extensive or intensive and either physical or chemical.
  Extensive properties, such as mass and volume, depend on the amount of matter that is being measured. Intensive properties, such as density and color, do not depend on the amount of matter.
  Both extensive and intensive properties are physical properties, which means they can be measured without changing the substance's chemical identity. For example, the freezing point of a substance is a physical property: when water freezes, it's still water (H_2O)—it's just in a different physical state.

- Chemical properties are those properties which are involved when matter undergoes a change to form new substances. A chemical property, meanwhile, is any of a material's properties that becomes evident during a chemical reaction; that is, any quality that can be established only by changing a substance's chemical identity. Chemical properties cannot be determined just by viewing or touching the substance; the substance's internal structure must be affected for its chemical properties to be investigated. The rusting of iron is a chemical property of iron since a new substance, iron rust, is formed.
  Here are some examples of chemical properties:

- Heat of combustion is the energy released when a compound undergoes complete combustion (burning) with oxygen. The symbol for the heat of combustion is ΔH_c.
- Chemical stability refers to whether a compound will react with water or air (chemically stable substances will not react). Hydrolysis and oxidation are two such reactions and are both chemical changes.
- Flammability refers to whether a compound will burn when exposed to flame. Again, burning is a chemical reaction—commonly a high-temperature reaction in the presence of oxygen.
- The preferred oxidation state is the lowest-energy oxidation state that a metal will undergo reactions in order to achieve (if another element is present to accept or donate electrons).

Physical and Chemical changes

Matter undergo two types of change;

Physical Change: This is a change which is easily reversed and in which no new substance is formed.

Examples of physical change include:

- The dissolution of common salt in water.i.e. Salt + water → Salt Solution
- Changes in the states of matter such as the melting of solids to liquids, the freezing of liquid to solids, the vaporization of liquids to gases, the liquefaction of gases to liquids and the sublimation of solids to vapours.
- The separation of mixtures by evaporation, distillation, fractional distillation, sublimation and crystallization.
- The magnetization and demagnetization of iron rods.

Chemical Change: This is a change which is not easily reversed and in which new substance is formed.

Examples of chemical change include:

- Burning of substance to ash
- The dissolution of metals and limestone in acids
- The rusting of iron.
- The addition of water to quicklime i.e. the slaking of lime.

- Fermentation and decay of substances.
- The changes in the electrochemical cell.

Difference between Physical and Chemical Changes

Physical Change	**Chemical Change**
1. It is easily reversible	It is not easily reversible
2. New substances are formed	New substances are formed
3. Very little amount of heat is involved	There are remarkable heat change
4. No change in mass	The new substances always have different masses

Assessment

1. The addition of water to calcium oxide leads to
   a. a physical change
   b. a chemical change
   c. the formation of mixture
   c. an endothermic change

2. Which of the following is an example of a chemical change
   a. dissolution of salt in water
   b. rusting of iron
   c. melting of ice
   d. separating a mixture by distillation

3. Which of the following is a physical change
   a. freezing ice-cream
   b. dissolving calcium in water
   c. burning kerosene
   d. exposing white phosphorous to air

4. A chemical reaction is always associated with
   a. an increase in the composition of one of the substance
   b. a change in volume of reactants

c. a change in nature of reactants
d. the formation of new substances

5. A heterogeneous mixture can be defines as any mixture
a. whose composition is uniform
b. whose composition is not uniform
c. formed by solids and liquids
d. of a solute and solvent

6. The dissolution of common salt in water is a physical change because
a. the salt can be obtained by crystallization
b. the salt can be recovered by evaporation of water
c. Heat is not generated during the mixture
d. the solution will not boil at 100°C

Answers

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

Week: 3

Topic: Elements

Element

An Element is a substance which cannot be split into simpler units by an ordinary chemical process. Most elements are not pure. The atoms are particles of elements and are represented by chemical symbols. A chemical symbol is a sign and it is made up of a letter or two letters to represent the atom of the element. Chemical symbols of atom of elements are derived from the names of the elements. There are 109 known elements. 90 of them occur naturally while the rest are artificially made in the laboratory. We use abbreviations to represent elements. e.g. Hydrogen as H, Nitrogen as N, Iron as Fe. Elements are grouped in form of a periodic table.

The following rules are used for deriving symbols:

- The first letter of the name is taken and this is written as a capital letter e.g. C for Carbon.
- The first and the second letters of the name is taken and first is written in capital letter while the second is in small letter e.g. Ca for Calcium, Al for Aluminium.
- Some first and the third letters of the name is taken and the first is written in capital letter while the third is in small letter e.g. Mg for Magnesium
- Some elements have Latin names, the symbols are derived from the Latin names e.g. The symbol for copper is Cu from Latin name is cuprum

There are six types of elements

1. the reactive metals
2. the transition metals
3. the lanthanides and actinides
4. the pure metals
5. the non-metals
6. the noble gases

More commonly, elements are grouped as non – metals, semi-metals or metalloids, and metals

A list of some elements is shown below:

Element	Symbol	Atomic Number	Atomic Mass	Number of Protons	Number of Neutrons
Hydrogen	H	1	1	1	0
Helium	He	2	4	2	2
Lithium	Li	3	7	3	4
Beryllium	Be	4	9	4	9
Boron	B	5	11	5	6
Carbon	C	6	12	6	6
Nitrogen	N	7	14	7	7
Oxygen	O	8	16	8	8
Fluorine	F	9	18	9	9
Neon	Ne	10	20	10	20
Sodium	Na	11	23	11	12
Magnesium	Mg	12	24	12	12
Aluminum	Al	13	27	13	14
Silicon	Si	14	28	14	14
Phosphorous	P	15	31	15	16
Sulphur	S	16	32	16	16
Chlorine	Cl	17	35	17	16
Argon	Ar	18	40	18	22
Potassium	K	19	39	19	20
Calcium	Ca	20	40	20	20

Iron	Fe	26	56	26	30
Copper	Cu	29	63.5	29	34.5
Zinc	Zn	30	65.5	30	35.5

Valency

Valency is defined as the combining power of any element. Valence describes how easily an atom or radical can combine with other chemical species. This is determined based on the number of electrons that would be added, lost or share if it reacts with other atoms. Valence is denoted using a positive or negative integer used to represent this binding capacity. For example, common valences of copper are 1 and 2.

Table of Element Valencies

Number	Element	Valency
1	Hydrogen	1
2	Helium	0
3	Lithium	1
4	Beryllium	2
5	Boron	3
6	Carbon	2 and 4
7	Nitrogen	3 and 5
8	Oxygen	2
9	Fluorine	1
10	Neon	0
11	Sodium	1
12	Magnesium	2
13	Aluminum	3
14	Silicon	4

15	Phosphorus	3 and 5
16	Sulphur	2, 4 and 6
17	Chlorine	1
18	Argon	0
19	Potassium	1
20	Calcium	2
21	Scandium	3
22	Titanium	2 and 4
23	Vanadium	3 and 5

Classification of Elements

Elements are classified as metals, metalloids and non-metals. Metals are those elements which ionize by electron loss. Examples are Calcium, Magnesium, Iron.

$Mg \rightarrow Mg^{2+} + 2e$

Nonmetals are those elements which ionize by electron gain. Examples are Chlorine, Oxygen, Nitrogen, Iodine.
$Cl + e \rightarrow Cl^{-}$

Differences between the Physical properties of Metals and Non Metals

METALS	NON-METALS
Metals are malleable, ductile and sonorous	Non – metals are not malleable, ductile and sonorous
Metals have great tensile strength	Non – metals are brittle or soft
Metals are lustrous	Non – metals are non-lustrous
Metals are good conductors of heat or electricity	Non – metals are non-conductors or poor conductors except graphite

Metals have high melting point and boiling point	Non-metals except carbon have low boiling and melting point
They have relatively high densities	They have low densities

Differences between Chemical Properties of Metals and Non-metals

They are reducing agents. They have the tendency to ionize by losing electrons to form cations	They are oxidizing agents. They tend to ionize by gaining electrons to form anions
They form basic oxides with oxygen	They form acidic oxides with oxygen
Some reactive metals can displace hydrogen from acids	They cannot displace hydrogen from acids
The generally tend to form ionic compounds	They tend to form covalent compounds

Some Compounds and their constituent elements

COMPOUND	COMPONENT ELEMENTS	FORMULA
Water	Hydrogen, Oxygen	H_2O
Sand	Silicon, Oxygen	SiO_2
Common salt	Sodium, Chlorine	$NaCl$
Sugar (sucrose)	Carbon, Hydrogen and Oxygen	$C_{12}H_{22}O_{11}$
Caustic Soda	Sodium, Hydrogen, Oxygen	$NaOH$
Ethanol	Carbon, Hydrogen, Oxygen	C_2H_5OH

Atoms

An atom is the smallest part of an element which can take part in a chemical reaction or the smallest part of an element that can ever exist and still possess the chemical properties of that element. An **atom** is the smallest constituent unit of ordinary matter that has the properties of a chemical element. Atoms are found to contain three types of particles: protons, neutrons and electrons.

It consists of centrally placed nucleus which is surrounded by a cloud of electrons. The nucleus is made of protons and neutrons. The proton and neutron each has a mass of one. Proton carries a positive charge while neutron has no charged. The electrons are very light and negatively charged and revolve round the nucleus in an orbital manner in order to neutralize the positive charge in the nucleus.

The atom is electrically neutral because the number of electrons revolving round the nucleus is equal to the number of protons in the nucleus.

The summary of the properties of sub-atomic particles are shown in the table below

Particle	Mass	Charge
Proton	1	+1
Neutron	1	No charge
Electron	1/1840	-1

Relationship between an Atom and an Element

An element is a basic substance that other things are composed from. Each individual element is made up of tiny, invisible particles called atoms. The atom is the smallest complete unit of an element. An atom is the smallest constituent unit of matter. Every atom is composed of a nucleus and one or more electrons which revolves around it continuously. An element is a chemical substance consisting of atoms having the same number of protons in their nuclei. Atoms and elements are related in that elements consist of one specific kind of atom. This means that the atom is the base unit of an element. For example, Hydrogen is an element composed of atoms that have one proton and one electron..

The arrangement of electrons in atoms

The ways in which the electrons are arranged outside and around the nucleus is called electronic configuration. The electrons are arranged in the orbits known as shells. The electrons revolve round the centrally placed nucleus in these shells.

There are seven shells in all and these are denoted by symbols K, L, M, N, O, P and Q. The nearest shell to the nucleus is K shell. It has the lowest energy and contains the smallest number of electrons which is two. It cannot hold more than two electrons.

To calculate the maximum number of electrons in a shell, the formula $2n^2$ is used.

N – Is the number of energy level of the shell.

When n = 2 for L shell, $2(2)^2 = 8$. The maximum number of electrons on the L shell is 8.

When n = 3 for M shell, $2(3)^2 = 18$. The maximum number of electrons on the M shell is 18 etc.

Atomic number and mass number

The atomic number of an element is the number of protons in one atom of that element. It is the number of protons in the nucleus of an atom. It is denoted by a letter Z.

The mass number of an element is denoted by A, which is the sum of the protons and the neutrons in the element. For example, sodium has a mass number of 23. The number of protons is 11 while neutrons is 12.

We can describe an atom of an element by writing its symbol together with its atomic number and mass number.

Using carbon as an example

Isotopy

Isotopy is a phenomenon whereby atoms of an element exhibit different mass numbers but have the same atomic number. This is due to differences in the number of neutrons present in these atoms. Such atoms are known as isotopes. Hence isotopes are atoms of the same elements with the same atomic numbers but different mass numbers. Isotopes of an element have slightly different physical properties but exhibit the same chemical properties. This is because neutrons contribute only to the mass of an atom, not its chemical behaviour.

Examples are as follow:

The mass spectrometer shows the relative abundance of the isotopes. The relative atomic masses of elements are the mean weight of the masses of the isotopes of the atoms of the elements. This is why they are not usually whole numbers. Given the mass numbers and relative abundance of the isotopes of an element, the relative atomic mass can be calculated.

Example 1: Chlorine exists in two isotopic form, ^{35}Cl and ^{37}Cl respectively. The relative abundance of 35Cl is 75% and 37Cl is 25%. Calculate the relative atomic mass of chlorine.

Solution:

Example 2: A naturally occurring sample of Lithium contains 7.42% of ^{6}Li and 92.58% of ^{7}Li. The relative mass of ^{6}Li is 6.015 and that of ^{7}Li is 7.016. Calculate the relative atomic mass of a naturally occurring sample of lithium.

Example 3: The relative atomic mass of an element is 21.5. The two isotopes of the element are ^{20}E occurring 25% in nature and ^{x}E occurring 75% in nature. Calculate x.

Solution:

Atomic mass = 20 x 25 + X x 75

100

Since relative atomic mass = 21.5, therefore

21.5 = 500 + 75x

100

X = 2150 – 500

75

X = 22

Assessment

1. The relative atomic mass of an element is 35. The two isotopes of the element are ^{20}E occurring 50% in nature and ^{x}E occurring 50% in nature. Calculate x.
2. What are the elements in the following compounds
   a. Tetraoxosulphate(vi) acid
   b. Methane
   c. Calcium hydroxide
   d. Galactose
   e. Nitrite
3. .... is defined as the combining power of any element

Week: 4

Topic: Elements (II)

Molecules

A molecule is the smallest particle of a substance that can normally exist alone and still retain the chemical properties of the substance be it an element or a compound. A *molecule* is the smallest particle in a chemical element or compound that has the chemical properties of that element or compound. *Molecules* are made up of atoms that are held together by chemical bonds. These bonds form as a result of the sharing or exchange of electrons among atoms. The number of atoms in each molecule is called **atomicity** of the element. Most gaseous elements like Oxygen and Chlorine are diatomic i.e the molecule consists of two atoms. Others like phosphorus and sulphur exists as polyatomic molecules. Helium is monoatomic – one molecule.

ELEMENT	FORMULA OF MOLECULE	ATOMICITY
Neon	Ne	1
Argon	Ar	1
Hydrogen	H_2	2
Nitrogen	N_2	2
Oxygen	O_2	2
Ozone	O_3	3
Phosphorus	P_4	4
Sulphur	S_8	8

The number of atoms in the molecule of a compound may be small or large e.g. Hydrogen Chloride molecule contains only 2 atoms while starch molecule contains thousands of atoms. Diatomic hydrogen is written as H_2 to show that it contains 2 atoms of hydrogen. A molecule of a compound contains whole numbers of atoms of the component elements.

Dalton's Atomic Theory

In 1808, John Dalton proposed the Atomic Theory which can be summarized below.

1. All elements are made up of small, indivisible particles called atoms.
2. Atoms can neither be created or destroyed
3. Atoms of the same element are alike in every aspect and differ from atoms of all other elements
4. When atoms combine with other atoms, they do in simple ratios
5. All chemical changes result from the combination or separation of atoms.

The atomic theory was partially supported by experimental evidences deducted from the Law of conservation of mass, the Law of Definite proportions etc.

Modifications of Dalton's Atomic Theory

1. *All elements are made up of small indivisible particles called atoms*. This statement was proven wrong by Rutherford's theory – The atom is built up of three main types of sub-particles, the proton, the neutron and the electron. It is not an indivisible solid piece
2. *Atoms can neither be created or destroyed*. The statement holds good for ordinary chemical reactions and its embodied in the Law of Conservation of Mass. During Nuclear fission, Uranium – 235 nucleus is broken into smaller units which form simpler atoms while a tremendous amount of heat is released.
3. *Atoms of the same element are alike in every aspect and differ from atoms of all other elements*. The discovery of isotopes makes this statement unacceptable. Chlorine has 2 different atoms or isotopes which differ in their neutron content and hence in their relative atomic number, they have same atomic number and different chemical activities.
4. *When atoms combine with other atoms, they do in simple ratios*. This statement is true for only inorganic compounds which usually contains a few atoms per molecule. Carbon however from very large organic molecules such as proteins, fats and starch which contain thousands of atoms.

Assessment

1. Atom is .....
   a. the smallest part of a substance that can take part in any chemical change
   b. the smallest part of a compound that can take part in any chemical change
   c. the smallest part of an element that can take part in any chemical change
   d. smallest part of a lattice that can take part in any chemical change
2. A molecule is the smallest particle of
   a. a matter that can exist in free state

b. an element that can exist in a free state
c. a radical that can exist in a free state
d. a lattice that can exist in a free state

3. $3NH_3$ is
   a. 3 moles of ammonium
   b. 3 moles of ammonia
   c. 6 moles of ammonia
   d. 6 moles of ammonium

4. Which one is heavier? 1 mole of Pbd_4, 1 mole of H_2 and 1 mole of $Pb(NO_3)_2$.
   a. Pbd_4
   b. H_2
   c. $Pb(NO_3)_2$
   d. None of the above

Answers

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

Week: 5

Topic: Particulate Nature of Matter

Introduction

Matter is defined as anything that occupies space and has mass. The particles of matter could be atoms, molecules or ions.

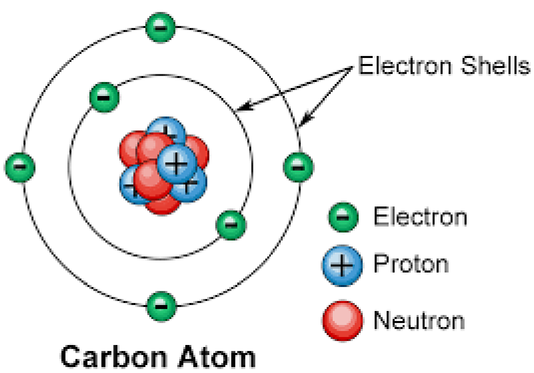
Ions

An ion is any atom or group of atoms which possesses an electric charge. There are two types of ions:

1. The positively charged ions or cations, examples are Na^+, Ca^+, Fe^{3+}, NH_4^+, etc.

2. The negatively charged ions or anions, examples are Cl^-, NO_3^-, OH^-, etc.

Atomic Structure



Atoms are made up of three sub-particles called electrons, protons and neutrons. The proton has positive (+ve) charge and a relative mass of 1 (using carbon -12 as standard). the electron has a negative (-ve) charge of the same magnitude as the positive charge on the proton, and negligible mass. The neutron has no charge but has a relative mass of 1. The characteristics of the neutron is the sum total of the proton and electron.

Proton is located in the Nucleus, relative charge +, relative mass 1

Electron is located outside the Nucleus, relative charge -, relative mass 0.0005

Neutron is located in the Nucleus, relative charge zero, relative mass 1.

Atomic Sub-particles

Discovery of Electrons

Electrons were discovered by Sir John Joseph Thomson in 1897 in his cathode ray experiment. He subjected residual gas to a high potential difference at a very low pressure. He observed rays travelling in straight lines from the cathode. He called the rays cathode rays. After many experiments of cathode-rays, J.J. Thomson demonstrated the ratio of mass to charge of cathode-rays which is 1.76×10^{11} coulombs per kg. He confirmed that cathode-rays are fundamental particles that have a negative charge. Cathode-rays became known as electrons. Robert Millikan in 1910, through oil-drop experiments, found the value of the charge – 1.6×10^{-19} coulom

In 1897, J J Thomson showed that when a potential difference of 5000V was applied across a glass tube containing a gas at a very low pressure of about 0.0001 atm, the tube began to glow. When the potential difference was increased to 15000V, a bright green glow appeared on the glass. Thomson was able to prove that the glow was due to some kind of rays which travelled in straight lines from the cathode. He called these cathode rays. Further experiments showed that cathode rays

1. travel in straight lines and cast shadows of opaque objects placed in their paths
2. are composed mainly of negatively charged particles
3. are capable of producing mechanical motion
4. are identical in nature

Thomson argued that these particles were fundamental and were present in all particles. He concluded that these particles must be the electrons proposed earlier to explain the conclution of an electric current. Thomson measured the ratio

$$\frac{\text{charge}}{\text{mass}}$$

which is the specific charge of the electron.

Discovery of Neutrons

Neutrons were discovered by James Chadwick in 1932 when he demonstrated that penetrating radiation incorporated beams of neutral particles. Neutrons are located in the nucleus with the protons. Along with protons, they make up almost all of the mass of the atom. The number of neutrons is called the neutron number and can be found by subtracting the proton number from the atomic mass number. The number of neutrons does not have to equal that of the protons.

Discovery of Protons

Since the atom is electrically neutral, there must exist inside the atom enough positively charged components to balance the negative charge of the electrons. He repeated the earlier experiments but use a discharge tube with central cathode which had a hole in it. He noticed a reddish glow in the opposite direction to the green glow and proved that the reddish glow was due to a positively charged ray.

Orbitals and Electronic Configuration

Orbital is a region or space where the probability of finding an electron is high while shell is an imaginary line on which electron revolves. Each shell is divided into orbitals. These orbitals are called s, p, d and f orbitals. Each shell contains its orbitals e.g. K shell contains only s orbital, while L shell contains s and p orbital, M shell contains s, p and d orbitals, N shell contains s, p, d and f orbital.

Each orbital has a maximum number of electrons it can hold. The sub-level 's' has a maximum of 2 electrons. The sub-level 'p' has a maximum of 6 electrons. The sub-level 'd' has a maximum of 10 electrons. The sub-level 'f' has a maximum of 14 electrons.

Orbital Types in Shell

Shell	Shell Number	Orbital Types
K	1	1s
L	2	2s, 2p
M	3	3s, 3p, 3d
N	4	4s, 4p, 4d, 4f

The sequence of filling up the orbitals with electrons is as follows:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4s etc.

Quantum Numbers

Each electron in an atom is described by four different quantum numbers. The first three (n, l, m) specify the particular orbital of interest, and the fourth (m_s) specifies how many electrons can occupy that orbital.

Principal Quantum Number (*n*): $n = 1, 2, 3, \ldots, \infty$
Specifies the energy of an electron and the size of the orbital. All orbitals that have the same value of *n* are said to be in the same shell (level). For a hydrogen atom with n=1, the electron is in its *ground state*; if the electron is in the n=2, it is in an *excited state*.

Subsidiary or Azimuthal Quantum Number (*l*): l = 0 to (n–1).
It specifies the shape of an orbital with a particular principal quantum number. The subsidiary quantum number divides the shells into smaller groups of orbitals called subshells (sublevels). The electrons with subsidiary quantum numbers 0, 1, 2 and 3 are usually referred to as the s, p, d, and f electrons respectively. The subshell with n=2 and l=1 is the 2*p* subshell; if n=3 and l=0, it is the 3*s* subshell, and so on. The value of *l* also has a slight effect on the energy of the subshell.

Magnetic Quantum Number (*m*): $m = -l, \ldots, 0, \ldots, +l$.
This number divides the subshell into individual orbitals which hold the electrons; there are $2l+1$ orbitals in each subshell. Thus the *s* subshell has only one orbital, the *p* subshell has three orbitals, and so on.

Spin Quantum Number (*s*): $s = +½$ or $-½$.

An electron can spin in only one of two directions (sometimes called *up* and *down*).

	s orbitals	*p orbitals*	*d orbitals*	*f orbitals*
L	0	1	2	3
M	0	-1, 0, +1	-2, -1, 0, +1, +2	-3, -2, -1, 0, +1, +2, +3
Number of orbitals in designated subshell	1	3	5	7

n	*l*	*M*	*Number of orbitals*	*Orbital Name*	*Number of electrons*
1	**0**	**0**	**1**	**1*s***	**2**
2	**0**	**0**	**1**	**2*s***	**2**
	1	**-1, 0, +1**	**3**	**2*p***	**6**
3	**0**	**0**	**1**	**3*s***	**2**
	1	**-1, 0, +1**	**3**	**3*p***	**6**

	2	-2, -1, 0, +1, +2	5	3*d*	10
4	0	0	1	4*s*	2
	1	-1, 0, +1	3	4*p*	6
	2	-2, -1, 0, +1, +2	5	4*d*	10
	3	-3, -2, -1, 0, +1, +2, +3	7	4*f*	14

Principles or Rules that Govern Electron Filling Into Orbitals

Pauli Exclusion Principle – This principle states that two electrons in the same orbital of an atom cannot have same values for all four quantum numbers. This implies that if two electrons are considered while they are may have the same values of n, L and m and they will differ in s because while one is n = 1, l = 0, m = 0 and s = +1/2 and the other will be n = 1, l = 0, m = 0 and s = -1/2. No two electrons in an atom have identical number.

Hund's Rule – Electrons occupy each orbital singly first before pairing takes place in a degenerate orbital. This implies that the most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins. An atom with fifteen electrons has its electrons arranged as $\mathbf{1S^2\ 2S^2\ 2P^2_x\ 2P^2_y\ 2P^2_z\ 3S^2\ 3P^1_x\ 3P^1_y\ 3P^1_z}$

Aufbau Principle – In the building up of atom, electrons enter into the orbital in order of increasing energy. This means that electrons are fed into the orbitals starting at the lowest energy level before fillinh the higher energy level.

The shapes of the s and p orbital – The electrons found in a given shell do not have the same amount of energy. This is because electrons move about the nucleus in different ways. The s orbital is spherical in shape while the p orbitals has its electrons moving about in 3 axes x, y and z. The 3 orbitals are represented as P_1, P_2, and P_3 and thus the P orbital has a dumb bell shape.

The way electrons are arranged in an atom is determined by the order in which the sub shells occur on a scale of increasing energy levels. This is so because the ground state the electrons will be found in the lowest energy levels available. In hydrogen, which has an atomic number of 1, the electron must occupy the 1S sub level. In case of hydrogen, there is only one electron in the subshell so we denote it as $1S^1$. Helium has an atomic number of 2, Its two electrons will accomodate the 1S orbital. The electronic configuration is written as $1S_2$.

The table below shows the modern electronic configuration of the first twenty elements.

Name	Atomic Number	Electron Configuration
Hydrogen	1	$1s^1$
Helium	2	$1s^2$
Lithium	3	$1s^2\ 2s^1$
Beryllium	4	$1s^2\ 2s^2$
Boron	5	$1s^2\ 2s^2 2p^1$
Carbon	6	$1s^2\ 2s^2 2p^2$
Nitrogen	7	$1s^2\ 2s^2 2p^3$
Oxygen	8	$1s^2\ 2s^2 2p^4$
Fluorine	9	$1s^2\ 2s^2 2p^5$
Neon	10	$1s^2\ 2s^2 2p^6$
Sodium	11	$1s^2\ 2s^2 2p^6 3s^1$
Magnesium	12	$1s^2\ 2s^2 2p^6 3s^2$
Aluminum	13	$1s^2\ 2s^2 2p^6 3s^2 3p^1$
Silicon	14	$1s^2\ 2s^2 2p^6 3s^2 3p^2$
Phosphorus	15	$1s^2\ 2s^2 2p^6 3s^2 3p^3$
Sulphur	16	$1s^2\ 2s^2 2p^6 3s^2 3p^4$
Chlorine	17	$1s^2\ 2s^2 2p^6 3s^2 3p^5$
Argon	18	$1s^2\ 2s^2 2p^6 3s^2 3p^6$
Potassium	19	$1s^2\ 2s^2 2p^6 3s^2 3p^6 4s^1$
Calcium	20	$1s^2\ 2s^2 2p^6 3s^2 3p^6 4s^2$

Electron Structure of the Atom

In Rutherford's model of atomic structure, the electrons moved in orbits arounf the nucleus and were held in their orbits by electrostatic attraction to the positively charged nucleus. This attraction varied inversely as the square of the distance between the nucleus and the electron.

Bohr's Model of the Atom

He made the following assumptions

1. Rutherford's model of the atom is correct
2. Each spectral line is caused by an electron
3. Electrons can exist only in circular orbits of definite quantum energy
4. An electron emits energy in the form of radiation when it moves from a higher to a lower permitted orbit – this produces a line in the atomic emission spectrum.
5. The difference in energy ΔE of the two orbits is related to the frequency of radiation given by Planck's equation
   $\Delta E = E_{n2} - E_{n1}$

Valency

This is the combining power of an element. The number of moles of hyrogen atoms that will combine with one mole of an element gives the valency of that element. The valency of an element is a property of its atomic structure. The valency allows you to construct the formulae of the compound from component elements.

GROUP	COMPOUND WITH HYDROGEN	THEIR VALENCY
I	LiH	1
II	BeH_2	2
III	BH_3	3
IV	CH_4	4
V	NH_3	3
VI	H_2O	2

VII	HF	1
O		O

Assessment

List the elements with the following configurations

1. $1s^2\ 2s^2 2p^5$
2. $1s^2\ 2s^2 2p^6 3s^2 3p^4$
3. $1s^2\ 2s^2 2p^6 3s^2 3p2$
4. $1s^2\ 2s^2 2p^1$
5. $1s^2\ 2s^2 2p3$

Week 6

Topic: The Relative Atomic Masses of Elements

Relative Atomic Mass

The mass of an atom is expressed as a ratio comparing the mass of one atom of any element with the mass of the hydrogen atom, the lightest known atom. The mass of an atom or the relative atomic mass carries no units, it is only a ratio. The relative atomic mass of Oxygen is 16 times that of Hydrogen while that of Sodium is 23 times that of Hydrogen. The hydrogen atom was then assigned a basic mass value of 1.

Relative atomic mass (symbol: A_r) is a dimensionless (number only) physical quantity. In its modern definition, it is the ratio of the average mass of atoms of an element in a given sample to one unified atomic mass unit. The unified atomic mass unit, symbol u, is defined being $^1/_{12}$ of the mass of a carbon-12 atom. The relative atomic mass of a naturally occurring element with isotopes can be defined as the weighted average of the masses of its isotopes compared with the mass of an atom of the carbon-*12* isotope whose mass is taken to be exactly *12*.

The relative atomic mass of an element is the number of times the average mass of one atom of that element is heavier than one twelfth the mass of one atom of carbon -12.

The relative atomic mass of Oxygen is given by
average mass of 1 atom of Oxygen
1/12 mass of 1 atom of carbn -12

Mass spectrometric studies show that the atoms of most elements exist in more than one form known as Isotopes. For instance, chlorine exists in two isotopes; one isotope ^{35}Cl, has a relative atomic mass of 35 and the other ^{37}Cl, has a relative atomic mass of 37. Isotopes of a given element exist in a constant ratio in nature.

Each isotope of an element has its own mass known as isotopic mass. The atomic mass of an element that exhibits isotopy s actually the weighted average isotopic mass of the isotopes of the element.
Every atom has its own unique relative atomic mass (RAM) based on a standard comparison or relative scale e.g. it has been based on hydrogen H = 1 amu and oxygen O = 16 amu in the past (amu = relative atomic mass unit).
The relative atomic mass scale is now based on an isotope of carbon, namely, carbon-12, nuclide symbol $^{12}_{6}C$, which is given the arbitrary value of 12.0000 amu by international agreement.

- The unit 'amu' is now being replaced by a lower case u, where u is the symbol for the unified atomic mass unit.
    - Therefore one atom of carbon, isotopic mass 12, equals 12 u, or,
    - 1 u = $^{1}/_{12}$th the mass of one atom of the carbon-12 isotope.
- Note that for the standard nuclide notation, $^{12}_{6}C$, the top left number is the mass number (12) and the bottom left number is the atomic/proton number (6).
- Since the relative atomic mass of an element is now based on the carbon-12 isotope it can now be defined as …
    - … relative atomic mass equals the average mass of all the atoms in an element compared to $^{1}/_{12}$th the mass of a carbon-12 atom (carbon-12 isotope).
    - Examples are shown in the Periodic Table diagram above.
    - Note
      (i) Because of the presence of neutrons in the nucleus, the relative atomic mass is usually at least double the atomic/proton number because there are usually more neutrons than protons in the nucleus (mass proton = 1, neutron = 1). Just scan the periodic table above and examine the pairs of numbers.
      You should also notice that generally speaking the numerical difference between the atomic/proton number and the relative atomic mass tends to increase with increasing atomic number. This has consequences for nuclear stability.
      (ii) For many calculation purposes, relative atomic masses are usually quoted and used at this academic level to zero or one decimal place eg.
        - e.g. hydrogen H = 1.008 or ~1; calcium Ca = 40.08 or ~40.0; chlorine Cl = 35.45 ~35.5, copper Cu = 63.55 or ~63.5/64, silver Ag = 107.9 or ~108 etc.
        - At Advanced level, values of relative atomic masses may be quoted to one or two decimal places.
          Many atomic masses are known to an accuracy of four decimal places, but for some elements, isotopic composition varies depending on the mineralogical source, so four decimal places isn't necessarily more accurate!
- In using the symbol A_r for RAM, you should bear in mind that the letter A on its own usually means the mass number of a particular isotope and amu is the acronym shorthand for atomic mass units.

- However there are complications due to isotopes and so very accurate atomic masses are never whole integer numbers.

- Isotopes are atoms of the same element with different masses due to different numbers of neutrons.

    - The very accurate relative atomic mass scale is based on a specific isotope of carbon, carbon-12, ${}^{12}C$ = 12.0000 units exactly, for most purposes C = 12 is used for simplicity.

    - For example ${}^{1}_{1}H$ hydrogen-1, ${}^{2}_{1}H$ hydrogen-2, and ${}^{3}_{1}H$ hydrogen-3, are the nuclide notation for the three isotopes of hydrogen, though the vast majority of hydrogen atoms have a mass of 1.

    - When their accurate isotopic masses, and their % abundance are taken into account the average accurate relative mass for hydrogen = 1.008, but for most purposes H = 1 is good enough!

- The strict definition of relative atomic mass (A_r) is that it equals the average mass of all the isotopic atoms present in the element compared to ${}^{1}/_{12}$th the mass of a carbon-12 atom (relative isotopic mass of 12.0000).

    - So, in calculating relative atomic mass you must take into account the different isotopic masses of the same elements, but also their % abundance in the element.

    - Therefore you need to know the percentage (%) of each isotope of an element in order to accurately calculate the element's relative atomic mass.

    - For approximate calculations of relative atomic mass you can just use the mass numbers of the isotopes, which are obviously all integers ('whole numbers'!) e.g. in the two calculations below.

    - To the nearest whole number, isotopic mass = mass number for a specific isotope.

Relative Atomic Mass of First 20 Elements

ATOMIC NUMBER	ELEMENT	ATOMIC MASS NUMBER
1	Hydrogen	1.008
2	Helium	4.0026
3	Lithium	6.939
4	Beryllium	9.0122
5	Boron	10.81
6	Carbon	12.01
7	Nitrogen	14.006
8	Oxygen	15.9994
9	Fluorine	18.9984
10	Neon	20.183
11	Sodium	22.9898
12	Magnesium	24.312
13	Aluminium	26.9812
14	Silicon	28.086
15	Phosphorus	20.987
16	Sulphur	32.06
17	Chlorine	35.453
18	Argon	39.948
19	Potassium	39.102
20	Calcium	40.08

Relative Molecular Mass

The relative molecular mass of an element is the sum of the relative atomic masses of all the atoms in one molecule of a substance. It s also referred to as formula mass.

Assessment

What do you understand by Relative Atomic Mass. Explain in your own words

Week 7

Topic: Compounds

Compounds

An element or compound is a pure substance because it cannot be separated into more than one component by physical methods.

A mixture consists of more than one element or compound. The components of a mixture can be separated by physical methods.

A compound is a substance which contains two or more elements chemically combined together. A compound is formed as a result of a chemical change. It is a new substance with entirely different properties from those of

- the substance from which it was formed
- the component elements

The component elements of a given compound are always present in a fixed ratio by mass. An example, water is a compound formed as a result of a chemical reaction between the component elements hydrogen and oxygen respectively in the ratio 2:1 respectively.
A compound is a substance formed when two or more chemical elements are chemically bonded together. Two types of chemical bonds common in compounds are covalent bonds and ionic bonds. The elements in any compound are always present in fixed ratios.

The constituents of mixtures can be elements or compounds or both. In samples of a given mixture, the constituents may be present in different proportions e.g different samples of cement contain variable proportions of calcium and aluminium trioxosilicate(IV). The constituents of mixture retain the individual identities because their physical and chemical properties are not changed by simple mixing.

1. Compounds consist of molecules formed from atoms of 2 or more different elements bound together chemically.
2. Compounds can be broken down into a simpler type of matter (elements) by chemical means; not by physical means
3. Compounds always contains the same ratio of component elements.
4. Compounds have properties different from their component elements
   e.g. the compound water (H_2O) is a liquid at room temperature and pressure and has

different chemical properties from those of the two elements, hydrogen (H_2) and oxygen (O_2), from which it is formed.

5. Compounds can be represented using chemical formulae.

Comparison of Mixture and Compound

MIXTURE	**COMPOUND**
It may be homogeneous or heterogeneous	It is always homogeneous
The constituents are not chemically bound together and can therefore be easily separated and recovered by physical means	The component elements are chemically bound together and cannot be separated by physical means
The constituents are added together in any ratio by mass. Hence a mixture cannot be represented by a chemical formula	The components are present in a fixed ratio by mass and can always be represented by a chemical formula
The properties of mixture are the sum of those of its individual constituents	The properties of a compound differ from those of its component elements

Examples of Compound and Mixtures

Compound

Water	Hydrogen and Oxygen	H_2O
Sand	Silicon and Oxygen	SiO_2
Limestone	Calcium, Carbon and Oxygen	$CaCO_3$
Caustic soda	Sodium, Hydrogen and Oxygen	$NaOH$
Common salt	Sodium and Chlorine	$NaCl$
Ethanol	Carbon, Hydrogen and Oxygen	C_2H_5OH

Mixture

Air	Oxygen, Carbon(IV) Oxide, Nitrogen, rare gases, dust, moisture

Soil	Sand, clay, humus, water, air, mineral salts
Urine	Urea, water, mineral salts
Palm wine	Water, sugar, alkanol, vitamins, yeast, proteins, fat
Coca cola	Water, sugar, carbon(IV) Oxide, coca cola concentrate
Blood	Water, proteins, fat, oil, sugar, vitamins, hormones, mineral salts, blood cells, haemoglobin

Relationship Between Elements, Compounds and Mixtures

- Elements can consist of either atoms or molecules but if molecules then those molecules are formed only from atoms of the same type (that is, atoms of the same element).
  For example, a molecule of oxygen consists of two atoms of oxygen and has the chemical formula O_2 where "O" is the chemical symbol of the element oxygen. Elements are either atoms or molecules, but only of one type of atom. Solid elements, e.g. metals, consist of many atoms of the element packed very densely together, e.g. copper, whose symbol is "Cu". Only a few elements exist naturally as a liquid at standard temperature and pressure. An example is the metal mercury, whose symbol is "Hg", which takes the form of many atoms that can move around each other easily but not away from the other atoms into the surrounding space, as in the case of gases.Many elements exist as a gas at standard temperature and pressure. In some cases, e.g. the inert or "noble" gases, the gas exists in the form of individual atoms, whereas in other cases the gas exists in the form of molecules (e.g. diatomic molecules – which means that each molecule is formed from two atoms of the the element, attached together), e.g. oxygen gas, whose formula is O2, nitrogen gas whose formula is N2 and hydrogen gas whose formula is H2. Elemental gases exist as either atoms or molecules according to the size and structure of the atoms of that element – which is usually explained later in school chemistry lessons
- Mixtures can consist of either atoms or molecules – but must include at least two different atoms or molecules.
  For example, a mixture of neon and argon gases would consist of atoms only because both neon and argon exist as atoms rather than as molecules. However, a mixture of oxygen and nitrogen gases would consist only of molecules because oxygen gas exists as oxygen molecules (O_2) and nitrogen gas exists as nitrogen molecules (N_2). A mixture of neon and nitrogen gases would consist of atoms of neon and molecules of nitrogen.
  Mixture is any combination of (different) : atoms + atoms, Or molecules + molecules or

atoms + molecules. That is, in the case of mixtures there must be at least two different types of atoms, or at least two different types of molecules, or at least one type of atom plus at least one type of molecule present. So, a mixture of gases may consist of molecules of the element oxygen (O2) plus molecules of the element nitrogen (N2) plus molecules of the compound carbon dioxide (CO2). Alternatively, a mixture of gases may consist of atoms of neon (Ne) and atoms of argon (Ar), and so on

- Compounds consist only of molecules (not individual atoms) and all the molecules of any one compound are the same. For example, methane gas has the chemical formula CH_4 because each molecule of methane is formed from one atom of carbon, whose chemical symbol is "C", and four atoms of hydrogen, whose chemical symbol is "H". Pure methane gas does not include any other atoms or molecules apart from the methane molecules described by the formula CH_4. Compounds consist of only one type of molecule: A molecule is the smallest part of a compound whose properties are those of the compound.

Formulae of Compounds

The molecules are particles of elements and compounds and they are represented by chemical formulae.

A chemical formula is the representation of molecules of elements and compounds by symbols. More importantly, it denotes the number of atoms of each element present in the compound. For example, the formula for Ferric oxide or Iron [III] oxide is Fe_2O_3, which implies that 2 atoms of Fe and 3 atoms of O are present in an electrically-neutral molecule of the compound.

Molecules of Elements and Atomicity

A molecule of an element may contain one atom or more.

- Molecule of element containing one atom of the element in the molecule is called a monoatomic molecule. The atomicity is 1. Examples are Helium (He), Neon (Ne), etc.
- Molecule of element containing two atoms of the element in the molecule is called a diatomic molecule. The atomicity is 2. Examples are Hydrogen (H_2), Oxygen (O_2), Nitrogen (N_2), Chlorine (Cl_2), etc.
- Molecule of element containing three atoms of the element in the molecule is called a triatomic molecule. The atomicity is 3. Example is Ozone (O_3)

Atomicity is the number of atoms in each molecule of an element.

Molecules of Compounds

When elements combine compounds are formed. To form compounds:

1. A metallic element can combine with a non-metallic element.
2. A non-metallic element can combine with a non-metallic element.

It must be noted that, a metallic element cannot combine with another metallic element but can be mixed to form alloy.

Compound may be classified as:

1. Binary Compounds: These compounds contain two elements only. Examples are NaCl, $MgCl_2$, Fe_2O_3, etc
2. Trinary Compounds: These compounds contain three elements only. Trinary compounds are derived from trinary acid. Examples are HNO_3, H_2SO_4, H_3PO_4, etc.

How to write a chemical formula

To write a chemical formula, one must know the symbols and valencies of the elements / radicals.

Example 1: Write the chemical formula for Calcium Tetraoxophosphate (V).
Valency of Calcium (Ca) = 2 ; Valency of Tetraoxophosphate (V) (PO_4) = 3.
Interchanging their valencies and writing as subscripts,
Formula for Calcium Tetraoxophosphate (V) is $Ca_3(PO_4)_2$.
Note that 3 calcium ion [Ca^{2+}] and 2 tetraoxophosphate ions [PO_4^{3-}] are present in an electrically-neutral molecule of calcium tetraoxophosphate (V) [$Ca_3(PO_4)_2$].

Example 2: Write the chemical formula for Zinc trioxocarbonate (IV).
Valency of Zinc (Zn) = 2 ; Valency of Trioxocarbonate (IV) (CO_3) = 2.
Interchanging their valencies and simplifying (on dividing by 2),
formula for Zinc carbonate is $ZnCO_3$.
Note that 1 zinc ion [Zn^{2+}] and 1 carbonate ion [CO_3^{2-}] are present in an electrically-neutral molecule of zinc carbonate [$ZnCO_3$].

Assessment

1. A compound is found to have a molecular mass of 90 atomic mass units and simplest formula of C_2H_5O. The molecular formula of the substance is:
   Use atomic masses of C = 12 amu, H = 1 amu, O = 16 amu

   A. $C_3H_6O_3$
   B. $C_4H_{26}O$

C. $C_4H_{10}O_2$
D. $C_5H_{14}O$

2. A substance of phosphorus (P) and oxygen (O) is found to have a mole ratio 0.4 moles of P for every mole of O.
The simplest formula for this substance is:A. PO_2
B. $P_{0.4}O$
C. P_5O_2
D. P_2O_5

3. Chemical formula of Zinc Hydroxide is

A. $Zn(OH)_2$
B. ZNO_2
C. $Zn_2(OH)_3$
D. ZnOH

4. Which sample contains the greatest number of molecules?
Atomic masses are given in parentheses
A. 1.0 g of CH_4 (16 amu)
B. 1.0 g of H_2O (18 amu)
C. 1.0 g of HNO_3 (63 amu)
D. 1.0 g of N_2O_4 (92 amu)

5. A sample of potassium chromate, $KCrO_4$, contains 40.3% K and 26.8% Cr. The mass percent of O in the sample would be:

A. 4 x 16 = 64
B. 40.3 + 26.8 = 67.1
C. 100 − (40.3 + 26.8) = 23.9
D. The mass of the sample is needed to finish the calculation.

6. How many grams of oxygen are in one mole of calcium carbonate, $CaCO_3$?
Atomic mass of O = 16 amu

A. 3 grams
B. 16 grams
C. 32 grams
D. 48 grams

7. The ionic compound containing Fe^{3+} and SO_4^{2-} would have the formula:

A. $FeSO_4$
B. Fe_2SO_4
C. $Fe_2(SO_4)_3$
D. $Fe_3(SO_4)_2$

Answers

1. C. $C_4H_{10}O_2$
2. D. P_2O_5
3. A.
4. A. 1.0 g of CH_4 (16 amu)
5. C. $100 - (40.3 + 26.8) = 23.9$
6. D. 48 grams
7. C. $Fe_2(SO_4)_3$

Week: 8

Topic: IUPAC NOMENCLATURE OF CHEMICAL COMPOUNDS

In chemical nomenclature, the IUPAC nomenclature of inorganic chemistry is a systematic method of naming inorganic chemical compounds, as recommended by the International Union of Pure and Applied Chemistry (IUPAC). It is published in Nomenclature of Inorganic Chemistry. Rules of naming are:

1. Single atom anions are named with an *-ide* suffix: for example, H^- is hydride.
2. Compounds with a positive ion (cation): The name of the compound is simply the cation's name (usually the same as the element's), followed by the anion. For example, NaCl is *sodium chloride*, and CaF_2 is *calcium fluoride*.
3. Cations which have taken on more than one positive charge are labeled with Roman numerals in parentheses. For example, Cu^+ is copper(I), Cu^{2+} is copper(II). An older, deprecated notation is to append *-ous* or *-ic* to the root of the Latin name to name ions with a lesser or greater charge. Under this naming convention, Cu^+ is cuprous and Cu^{2+} is cupric. For naming metal complexes see the page on complex (chemistry).
4. Oxyanions (polyatomic anions containing oxygen) are named with *-ite* or *-ate*, for a lesser or greater quantity of oxygen, respectively. For example, NO−2 is nitrite, while NO−3 is nitrate. If four oxyanions are possible, the prefixes *hypo-* and *per-* are used: hypochlorite is ClO^-, perchlorate is ClO−

5. The prefix *bi-* is a deprecated way of indicating the presence of a single hydrogen ion, as in "sodium bicarbonate" ($NaHCO_3$). The modern method specifically names the hydrogen atom. Thus, $NaHCO_3$ would be pronounced sodium hydrogen carbonate.

Positively charged ions are called cations and negatively charged ions are called anions. The cation is **always** named first. Ions can be metals or polyatomic ions. Therefore the name of the metal or positive polyatomic ion is followed by the name of the non-metal or negative polyatomic ion. The positive ion retains its element name whereas for a single non-metal anion the ending is changed to *-ide*. Example: sodium chloride, potassium oxide, or calcium carbonate.

Generally metals exhibit positive oxidation numbers while non-metals tend to have negative oxidation number. Some elements exhibit more than one valency. A few atoms of different elements may come together and react as one single unit. Such group is referred to as Radical.

Systems of Naming Compounds

Oxidation Number Concept – An oxidation number is a positive or negative number assigned to an atom according to a set of arbitrary rules.

Rules for Assigning Oxidation Number

1. The oxidation number of monoatomic ion is equal in magnitude and sign to its ionic charge. Example, the oxidation number of Bromine ion is Br^{-1}, is -1, that of Fe^{3+} is 3.

2. The oxidation number of hydrogen in a compound is always +1 except in metal hydrides for example NaH, where it is -1

3. Oxygen usually has oxidation number of -2 except in peroxides where it is -1

4. The oxidation number of uncombined elements is 0. Example, the oxidation number of N_2O_2, N_2 is 0

5. For any neutral compound, the sum of the oxidation number of the atoms of the compound must equal zero (0)

6. For a polymic ion, the sum of the oxidation numbers must equal the ionic charge of the ion.

- Binary Compounds: These compounds contain two elements only. Their names end with ide except water and ammonia. Examples are NaCl (sodium chloride), $MgCl_2$ (magnesium chloride), Fe_2O_3 (Iron(II) oxide)

- Trinary Compounds: These compounds contain three elements only. Trinary compounds are derived from trinary acid. Examples are HNO_3 (Hydrogen Trioxonitrate), H_2SO_4(Tetraoxosulphate(vi)acid), H_3PO_4, etc.

Oxidation number is the charge on the valency of an element, i.e. Oxidation state = Valency + Charge

The valency of metal carries a positive charge, e.g. Na^+ while the valency of non-metal carries the negative charge, e.g. Cl^-

The total oxidation state of any compound is zero. This is gotten by adding the oxidation number of all the elements in a compound. For example, zinc chloride $ZnCl_2$

Oxidation number of Zn = +2 and Cl = -1

Oxidation number of Zn + 2 x Oxidation number of Cl

= +2 + (2 x -1)

= +2 – 2 = 0

Some elements exhibit variable oxidation state, e.g. copper, lead, nitrogen and iron.

Example 1: Find the oxidation state of nitrogen in the following: (i) N_2O (ii) NO (iii) NH_3 (iv) NO_2 (v) NO_3

Solution:

(i) The sum of oxidation state of N_2O = 0

The oxidation number of oxygen = -2

$N_2O = 0$

$2 \times N + (-2) = 0$

$2N - 2 = 0$

$2N = +2$

$N = +2/2$

$N = +1$

The oxidation state of nitrogen in N_2O is +1 and the name of the compound is dinitrogen (I) oxide.

(ii) Oxidation state of nitrogen in NO

$NO = 0$

$N + (-2) = 0$

$N = 0 + 2$

$N = +2$

The oxidation state of nitrogen in NO is +2. Hence, the name is nitrogen (II) oxide.

(iii) Oxidation state of nitrogen in NH_3

$N + 3H = 0$

$N + 3 \times +1 = 0$

$N = 0 - 3$

$N = -3$

The oxidation state of nitrogen in NH_3 is -3. Hence, the name is ammonia.

(iv) Oxidation state of Nitrogen in NO_2

$N + 2O = 0$

$N + 2 \times -2 = 0$

$N = 0 + 4$

$N = +4$

The oxidation state of nitrogen in NO_2 is +4. Hence, the name is nitrogen (IV) oxide

(v) In a radical or an ion, the sum of the oxidation state is equal to its charge.

Oxidation state of nitrogen in NO_3^-

$N + 3O = -1$

$N - 2 \times 3 = -1$

$N - 6 = -1$

$N = -1 + 6$

$N = +5$

The oxidation state of nitrogen in NO3- is +5. Hence, the name is Trioxonitrate (V).

Example 2: Find the oxidation number of the manganese in potassium tetraoxomanganate (VII) $KMnO_4$

Solution: Oxidation state of potassium is +1

Oxidation state of oxygen is -2

kMnO4 = 0

$K + Mn + 4O = 0$

$+1 + Mn + 4 \times -2 = 0$

$+1 + Mn - 8 = 0$

$Mn - 7 = 0$

$Mn = 0 + 7$

$Mn = +7$

The oxidation number of the manganese in potassium tetraoxomanganate (VII) is +7.

Example 3: Find the oxidation number of the chromium in potassium heptaoxodichromate (VI) $K_2Cr_2O_7$.

Solution: Oxidation number of K = +1 and Oxidation number of oxygen = -2

$K_2Cr_2O_7 = 0$

$+1 \times 2 + Cr \times 2 + (-2 \times 7) = 0$

$+2 + 2Cr - 14 = 0$

$2Cr - 12 = 0$

$2Cr = 0 + 12$

$Cr = {}^{+12}/_{2}$

$Cr = +6$

The oxidation number of chromium in potassium heptaoxodichromate (VI) is +6

The IUPAC names of some salts:

$KClO_4$ = Potassium tetraoxochlorate (VII)

$Ca(NO_3)_2$ = Calcium trioxonitrate (V)

$CuSO_4$ = Copper (II) tetraoxosulphate (VI)

Na_2CO_3 = Sodium trioxocarbonate (IV)

$FeSO_4$ = Iron (II) tetraoxosulphate (VI)

$Fe(SO_4)_3$ = Iron (III) tetraoxosulphate (VI)

Example: Interpret the compounds represented by the formulae (i) $Ca(NO_3)_2$ (ii) $Fe_2(SO_4)_3$ (iii) KNO_3

Solution:

(i) $Ca(NO_3)_2$ represents 1 molecule of calcium trioxonitrate (V), containing:

1 atom of Calcium (Ca)

2 atoms of Nitrogen (N)

6 atoms of Oxygen (O)

(ii) $Fe_2(SO_4)_3$ represents 1 molecule of iron (III) tetraoxosulphate (VI), containing:

2 atoms of Iron (Fe)

3 atoms of Sulphur (S)

12 atoms of Oxygen (O)

(iii) KNO_3 represents 1 molecules of potassium trioxonitrate (V), containing:

1 atom of Potassium (K)

1 atom of Nitrogen (N)

3 atoms of Oxygen (O)

Naming Chemical Compounds

Naming chemical compounds are very important to identify the chemical compounds. Before naming the compound we should know the names and symbols for all elements like group I, and group II, halogen group, noble gases and some metals which are used frequently also we have to found what kind of compound it is.

- Chemical compounds are mainly divided in to two major types, such as ionic compound which contain metal and nonmetal, and covalent compound which contain only non metals.
- Ionic compounds are formed when metal give up its electrons to non metal. These ionic compounds can be easily recognized and naming them also very easy.
- Ionic compounds can further subdivided in to ionic compound without transition metal and ionic compound with transition metal.
- For naming ionic compounds without transition metal, first name the first element of the metal compound, simply name the metal name of the first element. Then name the second element of the non metal compound.
- The name of the non metal element should ending with suffix **'ide'**. For example consider Al_2O_3, the subscript of the compound indicates that how many atoms of elements present in that compound.
- These subscripts do not affect the name of the compound. So the first element name of the compound is named as aluminum.

Then name the second element of the compound is oxide, we drop the ending on oxygen and add **'ide'**.

Example :
$AlCl_3$ – Aluminum chloride
Na_2S – Sodium sulfide
K_2O – Potassium oxide

For naming ionic compound with transition metal is some what difficult then ionic compound without transition metal, since the transition metal can form more than one compound. A transition metal is an element with an atomic number of 21-30, 39-48, and 51-80 from the pediatric table. The transition metal in d and f blocks has more than one charge.

For example, the combination of iron and chlorine can form two different compounds, like $FeCl_2$ and $FeCl_3$. In such case we use Roman numeral to indicate the charge on the metal ions. For naming of ionic compound with a transition metal, specify the charge of the transition metal ions with Roman numeral.

Example :
Fe (Cl_2) – Iron (II) chloride
Fe (Cl_3) – Iron (III) chloride

The ions which are having two or more nonmetal atoms are called as polyatomic ions and they are covalently bonded together. Some of the polyatomic ion as follows.

The entire group of polyatomic ions has a positive or negative charge. Polyatomic ions are commonly forming two or more covalent compounds. The covalent compounds are formed from non-metals that share electrons. For example carbon and oxygen can form both carbon monoxide as well as carbon dioxide we cannot call both the compound as carbon oxide. To distinguish this covalent compound prefixes are used, these prefixes indicate how many atoms of each element are present in the compound.

Example :
CO_2 – Carbon dioxide
N_2S_3 – Dinitrogen trisulfide

Assessment

1. Interpret the compounds represented by the formulae (i) H_2SO_4 (ii) $Pb_2(NO_3)_2$ (iii) $Mg_2(SO_4)_3$
2. (a) BaO_2 (b) $(NH_4)_2MoO_4$ (c)$Na_3Co(NO_2)_6$ (d) CS_2

Answers

2. (a) If the oxidation number of the oxygen in BaO_2 were -2, the oxidation number of the barium would have to be +4. But elements in Group IIA can't form +4 ions. This compound must be barium peroxide, $[Ba^{2+}][O_2^{2-}]$. Barium therefore is +2 and oxygen is -1.

(b) $(NH_4)_2MoO_4$ contains the NH^{4+} ion, in which hydrogen is +1 and nitrogen is -3. Because there are two NH^{4+} ions, the other half of the compound must be an MoO_4^{2-} ion, in which molybdenum is -6 and oxygen is -2.

(c) Sodium is in the +1 oxidation state in all of its compounds. This compound therefore contains the $Co(NO_2)_6^{3-}$ complex ion. This complex ion contains six NO^{2-} ions in which the oxidation number of nitrogen is +3 and oxygen is -2. The oxidation state of the cobalt atom is therefore +3.

(d) The most elecronegative element in a compound always has a negative oxidation number. Since sulfur tends to form -2 ions, the oxidation number of the sulfur in CS_2, is -2 and the carbon is +4.

Week: 9

Topic: Mixture

Mixture

A mixture consists of more than one element or compound. The components of a mixture can be separated by physical methods. A **mixture** is a material system made up of two or more different substances which are mixed but are not combined chemically.
A **mixture** refers to the physical combination of two or more substances in which the identities are retained and are mixed in the form of solutions, suspensions, and colloids.

The constituents of mixtures can be elements or compounds or both. In samples of a given mixture, the constituents may be present in different proportions e.g different samples of cement contain variable proportions of calcium and aluminium trioxosilicate(IV). The constituents of mixture retain the individual identities because their physical and chemical properties are not changed by simple mixing.

Mixtures can be either homogeneous or heterogeneous. A homogeneous mixture is a type of mixture in which the composition is uniform and every part of the solution has the same properties. A heterogeneous mixture is a type of mixture in which the components can be seen, as there are two or more phases present. One example of a mixture is air. Air is a homogeneous mixture of the gaseous substances nitrogen, oxygen, and smaller amounts of other substances. Salt, sugar, and many other substances dissolve in water to form homogeneous mixtures. A homogeneous mixture in which there is both a solute and solvent present is also a solution. Mixtures can have any amounts of ingredients.

Examples of Mixtures

Mixture

Air	Oxygen, Carbon(IV) Oxide, Nitrogen, rare gases, dust, moisture
Soil	Sand, clay, humus, water, air, mineral salts
Urine	Urea, water, mineral salts
Palm wine	Water, sugar, alkanol, vitamins, yeast, proteins, fat
Coca cola	Water, sugar, carbon(IV) Oxide, coca cola concentrate
Blood	Water, proteins, fat, oil, sugar, vitamins, hormones, mineral salts, blood cells, haemoglobin

Relationship Between Elements, Compounds and Mixtures

- Elements can consist of either atoms or molecules but if molecules then those molecules are formed only from atoms of the same type (that is, atoms of the same element).
  For example, a molecule of oxygen consists of two atoms of oxygen and has the chemical formula O_2 where "O" is the chemical symbol of the element oxygen. Elements are either atoms or molecules, but only of one type of atom. Solid elements, e.g. metals, consist of many atoms of the element packed very densely together, e.g. copper, whose symbol is "Cu". Only a few elements exist naturally as a liquid at standard temperature and pressure. An example is the metal mercury, whose symbol is "Hg", which takes the form of many atoms that can move around each other easily but not away from the other atoms into the surrounding space, as in the case of gases.Many elements exist as a gas at standard temperature and pressure. In some cases, e.g. the inert or "noble" gases, the gas exists in the form of individual atoms, whereas in other cases the gas exists in the form of molecules (e.g. diatomic molecules – which means that each molecule is formed from two atoms of the the element, attached together), e.g. oxygen gas, whose formula is O_2, nitrogen gas whose formula is N_2 and hydrogen gas whose formula is H_2. Elemental gases exist as either atoms or molecules according to the size and structure of the atoms of that element – which is usually explained later in school chemistry lessons

- Mixtures can consist of either atoms or molecules – but must include at least two different atoms or molecules.
  For example, a mixture of neon and argon gases would consist of atoms only because both neon and argon exist as atoms rather than as molecules. However, a mixture of oxygen and nitrogen gases would consist only of molecules because oxygen gas exists as oxygen molecules (O_2) and nitrogen gas exists as nitrogen molecules (N_2). A mixture of neon and nitrogen gases would consist of atoms of neon and molecules of nitrogen.
  Mixture is any combination of (different) : atoms + atoms, or molecules + molecules or atoms + molecules. That is, in the case of mixtures there must be at least two different types of atoms, or at least two different types of molecules, or at least one type of atom plus at least one type of molecule present. So, a mixture of gases may consist of molecules of the element oxygen (O_2) plus molecules of the element nitrogen (N_2) plus molecules of the compound carbon dioxide (CO_2). Alternatively, a mixture of gases may consist of atoms of neon (Ne) and atoms of argon (Ar), and so on

- Compounds consist only of molecules (not individual atoms) and all the molecules of any one compound are the same. For example, methane gas has the chemical formula CH_4 because each molecule of methane is formed from one atom of carbon, whose chemical symbol is "C", and four atoms of hydrogen, whose chemical symbol is "H". Pure methane gas does not include any other atoms or molecules apart from the methane molecules described by the formula CH_4. Compounds consist of only one type of molecule: A molecule is the smallest part of a compound whose properties are those of the compound.

Assessment

Define the terms Mixture and Compounds. Give three differences between them.

Classify the following substances as an element, a mixture or a compound

- limestone
- diamond
- sand
- soil
- urine
- bronze
- sugar
- gold
- clay
- urea
- soap
- milk
- air
- iron

S.S.S 1

CHEMISTRY

SECOND TERM

TABLE OF CONTENT

Week 1

Topic: Mole Concept

Introduction

In chemistry, the term mole represents a pile or mass of atoms, molecules, ions or electrons. Just as common man measures quantity in terms of kilograms or dozens, a chemical scientist deals with a 'mole' of atoms, molecules, ions or electrons.

Mole is defined as the amount of a substance, which contains the same number of chemical units (atoms, molecules, ions or electrons) as there are atoms in exactly 12 grams of pure carbon-12. A mole is a unit which is used to express the amount of substance. In other words, it is defined as the amount of substance which contains Avogadro number of particles.

6.022×10^{23}, is called Avogadro's number (represented by N) named in the honour of an Italian scientist Amedeo Avogadro. The unit of mole is denoted as 'mol'.

$$\text{Number of moles (in mols)} = \frac{\text{mass of substance (in g)}}{\text{mass of one mole (in g mol}^{-1}\text{)}}$$

MOLE IN TERMS OF MASS

The mole is the amount of substance (Elements or compounds) which has a mass equal to its gram atomic mass or gram molecular mass.

Example 1: One mole of oxygen atoms = 16 g (One gm atomic mass)

Example 2: One mole of oxygen molecule = 32 g (One gm molecular mass)

MOLE IN TERMS OF NUMBER

One mole of substance contains Avogadro's number.

1 gm mole of hydrogen atom contains 6.022×10^{23} hydrogen atoms.

1 gm mole of Hydrogen molecule contains 6.022×10^{23} hydrogen molecule.

Molecular mass of water (H_2O) is 18

One Molecular mass of water (H_2O) is 18.

One mole of water (H_2O) contain 6.022×10^{23} molecules of water

Example 3: How many molecules are in 146 g HCl?

Solution:

$$146\,g \times \frac{1\text{ mol HCl}}{36.5\,g\text{ HCl}} \times \frac{6.02 \times 10^{23}\text{ molecules HCl}}{1\text{ mol HCl}} = 2.41 \times 10^{24}\text{ molecules HCl}$$

MOLE IN TERMS OF VOLUME

One mole of gas under standard temperature (273K) and Pressure 1 atm (STP) occupies 22.4 dm^3 of volume.

A mole of gaseous substance can also be defined as the amount of substance that can occupy the volume of 22.4 dm^3 at STP or 0.0224 m^3

1 mole = 6.022×10^{23} particles = 1gm

Molecular mass = 22.4 dm^3 at STP

Example 4: How many mole atoms of oxygen are there in 0.1 mol of carbon (IV) oxide (CO_2)?

Solution:

CO_2 represents 1 mole molecules of carbon (IV) oxide and it contains:

I mole atom of carbon (C)

2 mole of atoms of oxygen (O)

1 mole molecules of CO2 contains 2 mole atoms of O

0.1 mole molecules of CO2 contains (2 x 0.1)/1 mole atoms of O

= 0.2 mole atoms of Oxygen

Assessment

1. The equation for the formation of water is: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$. Calculate the number of moles of O_2 required to yield 6 moles of water.
2. Sodium combines with Oxygen as follows: $4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$. What is the mass of O_2 needed to burn 4.6g of Sodium.

Week 2

Topic: Mole Concept in Terms of Masses Numbers, Volumes of Reactants and Products

Molar Mass

The molar mass of any molecular compound is the mass in grams numerically equivalent to the sum of the atomic masses of the atoms in the molecular formula.
The mole is just a number; it can be used for atoms, molecules, ions, electrons, or anything else we wish to refer to. Because we know the formula of water is H_2O, for example, then we can say one mole of water molecules contains one mole of oxygen atoms and two moles of hydrogen atoms. One mole of hydrogen atoms has a mass of 1.008 g and 1 mol of oxygen atoms has a mass of 16.00 g, so 1 mol of water has a mass of (2 x 1.008 g) + 16.00 g = 18.02 g. The molar mass of water is 18.02 g/mol.

Example 1: The formula mass of methane, CH_4, calculates its molarmass.

Solution: C + H x 4

12.01 + (1.008 x 4) = 16.04 u and its molar mass is 16.04 g/mol.

Example 2: What is the molar mass of glucose, $C_6H_{12}O_6$?

Solution:

element	atomic mass	mass formula	Contribution
C	12.01 g/mol	x 6	= 72.06 g/mol
H	1.008 g/mol	x 12	= 12.10 g/mol
O	16.00 g/mol	x 6	= 96.00 g/mol
			total = 180.16 g/mol

The molar mass of ionic compounds can be calculated similarly, by adding together the atomic masses of all atoms in the formula to get the formula mass, and expressing the answer in units of g/mol.

Example 3: Calcium chloride, $CaCl_2$, has a molar mass of?

Solution: Ca + Cl x 2

40.08 + (2 x 35.45) = 110.98 g/mol.

Example 4: How many moles are there in 0.326 g of barium hydroxide?

Solution:

Molarmass of $Ba(OH)_2$ = 171

Mass of $Ba(OH)_2$ = 0.326g

Number of moles = 0.326/171 = 0.00191 mol

The molar volume

The volume occupied by one mole of a substance is called the molar volume for that substance. It may be easily calculated from the density of the substance:

Density = Mass/Volume

If one works in moles, this becomes

Density = Molar mass /Molar volume

The molar volume is of particular interest when dealing with gases. Since the density of a gas is greatly dependent on the temperature and pressure, the density should be measured at the standard temperature and pressure (STP), which is a temperature of 273K (0°C) and pressure of 1.0 atm or 760 mm of Hg or 760 torr or 76 cm of Hg.

It turns out that the molar volume (measured at STP) is nearly the same for all gases, and has the value of 22.4dm^3

To put it in another way, one mole of any gas at STP will occupy a volume of 22.4 dm^3

Percentage Composition of Compounds

Assume 1 mole of the compound

Percent = mass of the element / molar mass of compound x 100

Using the molar masses of the elements and the compound, we can express the composition in terms of mass percentage of the elements. For example, carbon dioxide has a formula weight of 44.01 u, made up of 12.01 u for the average mass of 1 carbon atom and 32.00 u for 2 oxygen atoms. One mole of CO_2 has a mass of 44.01 g made up of 12.01 g of carbon and 32.00 g of oxygen. The composition of carbon dioxide is calculated as follows:

12.01g carbon / 44.00 carbon dioxide x 100 = 27.30% carbon

32.00g oxygen / 44.00g carbon dioxide x 100 = 72.70% oxygen

The formula for a compound, and its composition expressed as percentage by mass, are fixed and unchanging properties of the compound. Any pure sample of carbon dioxide is 72.70 % oxygen by mass.

This property allows us to relate the amount of an element in a compound to the mass of the compound.

Example 1: Calculate the number of moles of iron in 2.98 g of iron (III) oxide.

Solution:

First we need the formula for iron (III) oxide. Since iron (III) is Fe^{3+} and oxide is O^{2-} the formula must be Fe_2O_3. Calculate the molar mass as (2 x 55.85) + (3 x 16.00) = 111.7 + 48.00 = 159.7 g/mol.

The percentage by mass of iron in iron (III) oxide can then be calculated

= 111.7 g iron / 159.7 g iron (III) oxide x 100 = 69.94% iron

This percentage by mass can be used to calculate the mass of iron in the sample

= 2.98 g iron (III) oxide x 69.94/100 = 2.08g iron

Finally, the mass of iron can be converted to an amount in moles using the molar mass

= 2.08 g iron / 55.85 g/mol = 3.73×10^{-2} mol

The mole of iron in 2.98 g of Fe_2O_3 is 3.73×10^{-2} mol.

Example 2: Calculate the average atomic weight of silicon having 92.2% of Si-28 isotope of relative mass 27.98 amu, 4.7% of Si-29 isotope of relative mass 28.98 amu and 3.1% of Si-30 isotope of relative mass 29.97 amu.

Solution:

Data given: 92.20% of silicon of mass 27.98 amu

4.7% of silicon of mass 28.98 amu

3.1 % of silicon of mass 29.97 amu

Average atomic weight

= 25.80 + 1.36 + 0.93

= 28.09 amu

Average atomic weight of silicon is 28.09 amu

Example 3: Calculate the % of copper in copper sulphate, $CuSO_4$

Solution:

Relative atomic masses: Cu = 64, S = 32 and O = 16

Relative formula mass = 64 + 32 + (4×16) = 160

Only one copper atom of relative atomic mass 64

% Cu = 64 x 100 / 160 = 40% copper by mass in the compound

To calculate the % of the other elements in the compound

% sulfur = (32/160) x 100 = 20% S

% oxygen = (64/160) x 100 = 40% O

Also note that if you haven't made any errors, they should add up to 100% !

Example 4: Calculate the % of oxygen in aluminium sulphate, $Al_2(SO_4)_3$

Solution:

Relative atomic masses: Al = 27, S = 32 and O = 16

Relative formula mass = 2×27 + 3(32 + 4×16) = 342

Giving a total mass of oxygen in the formula of 12 x 16 = 192

% O = 192 x 100 / 342 = 56.1% oxygen by mass in aluminium sulphate

Example 5: Calculate the % of water in hydrated magnesium sulphate $MgSO_4.7H_2O$

Solution:

Relative atomic masses: Mg = 24, S = 32, O = 16 and H = 1

Relative formula mass = 24 + 32 + (4 x 16) + [7 x (1 + 1 + 16)] = 246

7 x 18 = 126 is the mass of water

So % water = 126 x 100 / 246 = 51.2 % H_2O

Empirical and Molecular Formulae

Molecular Formula is a formula indicating the actual number of atoms of each element making up a molecule. The molecular formula must accurately state the exact number of atoms of all of the elements in one molecule of the substance.

Empirical formula is the formula giving the simplest ratio between the atoms of the elements present in a compound. You must find the ratio of atom to atom in a molecule, and then reduce it.

Facts on Empirical and Molecular Formulae

- Empirical Formula of a compound shows the ratio of elements present in a compound.
- Molecular Formula of a compound shows how many atoms of each element are present in a molecule of the compound.
- The empirical formula mass of a compound refers to the sum of the atomic masses of the elements present in the empirical formula.
- The Molecular Mass (formula mass, formula weight or molecular weight) of a compound is a multiple of the empirical formula mass.
  MM = n x empirical formula mass
- Empirical Formula can be calculated from the percentage (or percent) composition of a compound.

Creating an Empirical Formula From Mass or Percent Composition

From Percent Composition:

- Step 1: Create a chart with six columns and a number of rows equal to the number of elements in the compound.
- Step 2: Write the elements in the first column.
- Step 3: Write the percent composition of each element in the second column.
- Step 4: Using the percent composition as the mass, divide each by the molecular mass of the respective element.
- Step 5: Divide each of those numbers by the smallest of the numbers in that column to reduce the ratio. If one or more numbers in the ratio is still distant from a whole number, try multiplying the entire ratio by a whole number. If the number is close, then round it to the nearest whole number. Each one of the numbers is the subscript for the corresponding element.

Example 1: Determine the empirical formula for a compound containing 74.0% carbon, 8.65% hydrogen, and 17.3% nitrogen by mass.

C	74.0%	74.0 g / 12.0 g	6.16 / 1.24	4.96	5

H	8.65%	8.65 g / 1.01 g	8.56 / 1.24	6.90	7
N	17.3%	17.3 g / 14.0 g	1.24 / 1.24	1	1

Therefore the empirical formula is C_5H_7N

From Mass:

Follow the same procedure as fir Percent Composition but first divide the mass of each element by the mass of the total sample. The quotient is the percent composition.

Example 2: 9.2 g sample of a compound is 2.8 g of nitrogen and 6.4 g of oxygen. Find the empirical formula of the compound.

N	2.8 g / 9.2 g	30%	30 g / 14 g	2.143 g / 2.143 g	1	1
O	6.4 g / 9.2 g	70%	70 g / 16 g	4.375 g / 2.143 g	2.04	2

- Therefore, the empirical formula is NO_2.

Examples 3: If carbon and hydrogen are present in a compound in a ratio of 1:2 and the molecular mass of the compound is 28.0 g/mol.

Solution: The empirical formula for the compound is CH_2

The empirical formula mass of this compound is: 12.0 + (2 x 1.0) = 14.0 g/mol

Since the molecular mass of the compound is 28.0 g/mol then we can find the molecular formula for the compound.
MM = n x empirical formula mass
28.0 = n x 14.0
n = 2
So the molecular formula for the compound is 2 x empirical formula, i.e., 2 x (CH_2) which is C_2H_4 – Ethene

Example 4: A compound containing 92.3 weight percent of carbon and 7.7 weight percent of H. What is the empirical formula?

Solution:
Assume that you have 100 g of the compound, then you have 92.3 g of carbon and 7.7 g of hydrogen. Thus the mole ratio of C to H should be

C	H
92.3/12	7.7/1.08

7.69/7.13 7.13/7.13

1.07 1.0

Thus, the empirical formula is CH.

A compound with an empirical formula of CH has a molecular weight of 78 g/mol. What is the molecular formula?

Solution:
The formula weight of CH is 13.0.
Since 78/13 = 6,
The molecular formula is C_6H_6, the formula for benzene.

Assessment

1. How many atoms of C are contained in 1 mole of CO2 ?
2. How many atoms of O are contained in 1 mole of CO2 ?
3. Find mass of 1 molecule C_2H_6. (C=12, H=1)
4. Calculation of number of particles when the mass of the substance is given:

Number of particles = Avogadro number x given mass / gram molecular mass
Calculate the number of molecules in 11g of CO2

Answers

1. 1 mole CO2 × 1 mole C
   1 mole CO2
   × 6.02×1023atoms C
   1 mole C = 6.02×1023atoms of C
2. 1 mole CO2 × 2 moles O
   1 mole CO2
   × 6.02×1023atoms O
   1 mole O =1.20 ×1024 atoms of O
3. C_2H_6=2.12+6.1=306,02×10^{23} C_2H_6 molecule is 30 g

1 C_2H_6 molecule is ? g

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1 C_2H_6 molecule is =?=5.10^{-23} g

4. Solution: gram molecular mass of CO_2 = 44g
   44g of CO_2

= 6.023×10^{23}molecules
1 g of CO_2

= ($6.023 \times 10^{23} \div 44$ g) molecules 11g of CO_2= ($6.023 \times 10^{23} \div 44$ g) x 11 =

1.51×10^{23} molecules

Week 3

Topic: Writing and Balancing of Chemical Equation

Chemical Equation

A chemical equation is a concise shorthand expression which represents the relative amount of reactants and products involved in a chemical reaction.

$CO_2 + H_2O$ ——> H_2CO_3

- The first step in writing a chemical equation is the word equation. It is composed of the names of the materials that enter into chemical reaction, the reactants, and the names of the materials that result from the reaction, the products.
- The second step in writing a chemical equation is a skeleton equation. This equation includes the chemical formulas and symbols for all the reactants and products.
- The third step is a balanced equation. This equation is similar to a skeleton equation, but it also includes coefficients placed directly in front of the chemical formulas and symbols. The coefficients of a balanced equation indicate the number of units of each substance involved.

Hydrogen burns in oxygen to form water. The reactants are hydrogen and oxygen. The product is water. The word equation for this reaction is: hydrogen + oxygen ——-> water

It is read, "hydrogen plus oxygen yield water." Since hydrogen and oxygen are diatomic gases, the H_2 and O_2 represent one molecule of hydrogen and one molecule of oxygen. The compound, water, is represented by the formula H_2O. The skeleton equation is: $H_2 + O_2$ ——-> H_2O

The Balanced Equation

The skeleton equation indicates that two hydrogen atoms react with two oxygen atoms on the reactant side of the equation. On the product side, two hydrogen atoms are bonded to one oxygen atom. The equation is deficient by one oxygen atom on the product side. Balance the equation by putting coefficients directly in front of any of the reactants or products. A coefficient should be a whole number. Never change the subscripts. Changing the subscript changes the chemical formula of the compound. Place a 2 in front of the H_2O so that there are two oxygen atoms on each side of the equation.

$H_2 + O_2$ ——–> $2H_2O$

The 2 placed in front of the H_2O changes the balance of the hydrogen atoms. Correct this imbalance by placing another 2 in front of the H_2. The equation is now balanced.

$2H_2 + O_2$ ——–> $2H_2O$

Symbols are used in an equation to indicate the physical state of each substance. The symbols used (placed in parenthesis) are (g) for gas, (g) for liquid, (l) for solid, (s) and (aq) for aqueous (water) solution.

$2H_2(g) + O_2(g)$ ——–> $2H_2O(l)$

The equation indicates:

- Two molecules of hydrogen react with one molecule of oxygen to give two molecules of water
- A quantitative relationship between the reactants and products. This relationship is the basis for working all quantitative problems

Steps in Writing a Balanced Equation

- Step 1: In writing a chemical reaction, the reactants and the products are written down. If the products are not known, they can be predicted in many cases.
- Step 2: The formulas of each substance must be correct. The diatomic gases are hydrogen, H_2, nitrogen, N_2, oxygen, O_2, fluorine, F_2; chlorine, Cl_2, bromine, Br_2; and iodine, I_2.
- Step 3: The equation is then balanced. The balancing is done by putting a coefficient before compounds until the atoms of each element on one side of the equation equal the number of atoms of that element on the other side of the equation.

Example 1: Sodium reacts with water to produce a metallic hydroxide and hydrogen gas. Write a balanced equation for the reaction.

Solving Process:

Step 1: Write the word equation: sodium + water —–> sodium hydroxide + hydrogen

Step 2: Write a skeleton equation. Since hydrogen is a diatomic gas, its formula is H_2. The formula for water may be written as HOH if this makes it easier to balance the equation.

Na + HOH ——> NaOH + H_2

Step 3: Balance the equation.

The metallic element sodium is balanced. One atom of sodium is on each side of the equation. There is one hydrogen atom on the reactant side (the H in OH has been accounted for) and 2 hydrogen atoms on the product side. Place a 2 in front of the HOH to balance the hydrogen atoms.

$Na + 2HOH$ ——> $NaOH + H_2$

There are now 2OH on the left and 1 on the right. Place a 2 in front of the NaOH to give the same number of OH on each side.

$Na + 2HOH$ ——> $2NaOH + H_2$

Put a 2 in front of the sodium metal. The balanced equation reads

$2Na(c) + 2HOH(l)$ —–> $2NaOH(aq) + H_2(g)$

Example 2: Iron(III) oxide reacts with carbon monoxide to give iron and carbon dioxide. What is the balanced equation?

Solving Process:

Step 1:Write the word equation: iron(III) oxide + carbon monoxide —–> iron + carbon dioxide

Step 2: Write the formulas and symbols of all reactants and products.

$Fe_2O_3 + CO$ —–> $Fe + CO_2$

Step 3: Balance the iron atoms

$Fe_2O_3 + CO$ —–> $2Fe + CO_2$

The carbon atoms are already balanced. Visual inspection indicates that a 2 in front of the CO, to balance the oxygen atoms changes the carbon atom balance. A 2 in front of carbon monoxide results in an odd number of oxygen atoms on the reactant side. This change will not work because any number placed in front of carbon dioxide will always give an even number of oxygen atoms on the product side. A 3 in front of CO will give an even number of oxygen atoms on the reactant side. A 3 as a coefficient for CO, will balance the oxygen atoms and the carbon atoms. The balanced equation is

$Fe_2O_3(s) + 3CO(g)$ —–> $2Fe(s) + 3CO_2(g)$

Example 3: Calcium hydroxide reacts with phosphoric acid to yield calcium phosphate and water. Determine the balanced equation

Solving Process:

Step 1: calcium hydroxide + phosphoric acid —–> calcium phosphate + water

Step 2: $Ca(OH)_2 + H_3PO_4$ ——> $Ca_3(PO_4)_3 + HOH$

Step 3: Balance the calcium atoms by placing a 3 in front of the $Ca(OH)_2$

$3Ca(OH)_2 + H_3PO_4$ —–> $Ca_3(PO_4)_2 + HOH$

Since the phosphate and hydroxide ions are on both sides of the equation they can be balanced as units. Place a 2 in front of the phosphoric acid to balance the PO_4, group. Place a 6 in front of the HOH to balance the OH group.

$3Ca(OH)_2(s) + 2H_3PO_4(l)$ ——> $Ca_3(PO_4)_2(s) + 6HOH(l)$

Visual inspection shows that the hydrogen in H_3PO_4 and HOH is balanced

We have used only the smallest whole number coefficients. In balancing equations, you may sometimes obtain multiples of the smallest coefficients. If so, reduce the coefficients to the smallest whole numbers.

Assessment

Balance the following equations

1. $FeCl_3 + MgO$ —> $Fe_2O_3 + MgCl_2$
2. $Li + H_3PO_4$ —> $H_2 + Li_3PO_4$
3. $Fe + HC_2H_3O_2$ —> $Fe(C_2H_3O_2)_3 + H_2$
4. $HCl(aq) + MnO_2(s)$ —> $MnCl_2(aq) + Cl_2(g) + H_2O(\ell)$
5. $CO_2 + S_8$ —> $CS_2 + SO_2$

Answers

1. Balance the Cl (note that 2 x 3 = 3 x 2):
   $2FeCl_3 + MgO$ —> $Fe_2O_3 + 3MgCl_2$
   The Fe also gets balanced in this step.
   Pick either the O or the Mg to balance next:
   $2FeCl_3 + 3MgO$ —> $Fe_2O_3 + 3MgCl_2$
   The other element (Mg or O, depending on which one you picked) also gets balanced in this step.

2. Balance the Li:
   $3Li + H_3PO_4$ —> $H_2 + Li_3PO_4$
   Now, look at the hydrogens. See how the H comes only in groups of 3 on the left and only in groups of 2 on the right?
   Do this: $3Li + 2H_3PO_4$ —> $3H_2 + Li_3PO_4$
   Remember 2 x 3 = 6 and 3 x 2 = 6. It shows up a lot in balancing problems (if you haven't already figured that out!).

Balance the phosphate as a group: $3Li + 2H_3PO_4 \longrightarrow 3H_2 + 2Li_3PO_4$
Oops, that messed up the lithium, so we fix it: $6Li + 2H_3PO_4 \longrightarrow 3H_2 + 2Li_3PO_4$

3. Balance the acetate:
$Fe + 3HC_2H_3O_2 \longrightarrow Fe(C_2H_3O_2)_3 + H_2$
Balance the hydrogen:
$Fe + 3HC_2H_3O_2 \longrightarrow Fe(C_2H_3O_2)_3 + (3/2)H_2$
Clear the fraction:
$2Fe + 6HC_2H_3O_2 \longrightarrow 2Fe(C_2H_3O_2)_3 + 3H_2$

4. Balance the chlorine:
$4HCl\ (aq) + MnO_2(s) \longrightarrow MnCl_2(aq) + Cl_2(g) + H_2O\ (\ell)$
Balance the hydrogen:
$4HCl\ (aq) + MnO_2(s) \longrightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O\ (\ell)$

5. The only thing not balanced already is the S:
$CO_2 + {}^3/_8 S_8 \longrightarrow CS_2 + SO_2$
Most of the time the fraction used to balance is something with a 2 in the denominator: ${}^1/_2$ or ${}^5/_2$ or ${}^{13}/_2$, for example. Not too often does one see ${}^3/_8$. Pretty tricky!
Multiply through by 8:
$8CO_2 + 3S_8 \longrightarrow 8CS_2 + 8SO_2$

Week 4 & 5

Topic: Law of Chemical Combinations

Introduction

Compounds are formed by chemical combination of reactants (atoms or molecules) which may be solid, liquid or gaseous. Chemical combination occurs in definite proportion by weight or by volume. Based on various experiments performed by different scientists, the laws of chemical combinations were formulated. These laws laid the foundation of stoichiometry, a branch of chemistry in which quantitative relationship between masses of reactants and products is established. The study of these laws led to the development of a theory concerning the nature of matter. These laws summarize the experimental results that are obtained about the masses of the reactants and products in a very large number of chemical reactions. There are four laws of chemical combination which describe the general features of a chemical change.

The various laws are:

1. Law of conservation of mass
2. Law of definite proportions
3. Law of multiple proportions
4. Law of reciprocal proportions

Law of Conservation of Mass

This law states that the total mass of the reactants is equal to the total mass of the products in every chemical reaction. The total mass of material present after a chemical reaction is the same as before the reaction. This Law was discovered by Antoine Lavoisier in about 1789. Whenever a chemical change occurs, the total mass of products is the same as the total mass of reactants. Alternatively the law can be stated as "the total mass of substances taking part in a chemical reaction remains the same throughout the change." or " matter is neither created nor destroyed during chemical reactions but changes from one form to another"

Experimental verification of the law of conservation of mass

Experiment 1:

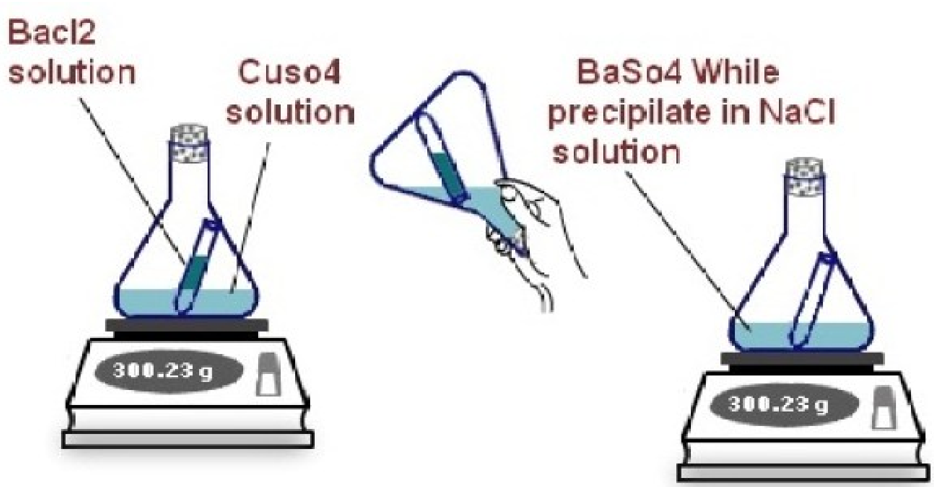
This law can be verified by the study of any chemical reaction. In the laboratory, it can easily be verified by the study of the following reaction.

$Ag\ NO_3 + NaCl \rightarrow AgCl \downarrow + NaNO_3$

When a solution of silver nitrate ($AgNO_3$) is treated with a solution of sodium chloride, a white precipitate of silver chloride ($AgCl$) is obtained along with a solution of sodium nitrate ($NaNO_3$). If the law is true, the total mass of $AgNO_3$ and $NaCl$ should be the same as the total mass of $AgCl$ precipitate and $NaNO_3$ solution. The experiment is done in a specially designed H shaped tube called Landolt's tube. Sodium chloride solution is taken in one limb of the tube while silver nitrate solution is taken in the other limb. Both the limbs are now sealed and tube is weighed. Now the tube is inverted so that the solutions can mix up together and react chemically. The reaction takes place as mentioned above and a precipitate of silver chloride is obtained. The tube is again weighed. The mass of the tube is found to be exactly the same as the mass obtained before inverting the tube. This experiment clearly shows that the law of conservation of mass is true

Experiment 2

The law can also be verified by the below mentioned experimental setup. Take a conical flask and a test tube. Take 10ml barium chloride solution in the conical flask and 10ml copper sulphate solution in the test tube. Measure the initial mass of reactants and note it.
Now mix the solutions together. Copper sulphate reacts with barium chloride to give a white precipitate of barium sulphate. Then take the weight of the products formed. We can observe that the mass of reactants before the reaction and the products formed after the reaction will be equal.



As a result of numerous experiments conducted on systems in closed containers, so as to retain any gases present, Lavoisier was able to unambiguously verify his assumptions and formally state the Law of Conservation of Mass. The law of conservation of mass states that matter can neither be created nor destroyed in the course of chemical reaction.

Note: When an energy difference occurs during a reaction, minute amounts of mass are either gained or lost. Mass is either converted to energy or energy is converted to mass. The energy-mass equivalence was first postulated by Einstein in his famous formula; $E = mc^2$.
While these mass differences are not detectable by the chemist, they are important in

nuclear reactions. Therefore, law of conservation of mass is applicable in all chemical reactions it is not applicable in the case of nuclear reaction where a fraction of the mass is converted to energy.

Example 1: In an experiment 5.0g of $CaCO_3$ on heating gave 2.8 g of CaO and 2.2 g of CO_2. Show that these results are in accordance to the law of conservation of mass.

Solution:

$CaCO_3$ ——> CaO + CO_2

Weight of CaCO3 = 5.0 gms

Weight of CaO = 2.8 gms

Weight of CO2 = 2.2 gms

Total weight of reactant = Total weight of products.

5.0 = 2.8 + 2.2

5.0 = 5.0

Since, the mass of the reactants are equal to the mass of the product, these results are in accordance to the laws of conservation of mass.

Example 2: In an experiment, 48 gms of magnesium combines with 32 gms of oxygen to form 80 gms of magnesium oxide. Show that this reaction illustrates the Law of Conservation of Mass. [Hint: 2Mg + O2 → MgO, Atomic mass of Mg = 24 and O = 16].

Solution:

2 Mg + O_2 → 2MgO

Magnesium + oxygen →Magnesium oxide

Weight of Magnesium = 48 gms

Weight of oxygen = 32 gms

Weight of magnesium oxide = 80 gms

∴ Total weight of reactants = Total weight of products

48 + 32 = 80

80 gms = 80 gm.

So these results are in accordance to the laws of conservation of mass.

Law of Definite Proportions

A chemical compound, no matter what its origin or its method of preparation, always has the same composition; i.e., the same proportions by mass of constituent elements. This Law, sometimes known as the Law of Definite Composition, was first enunciated by Joseph Proust in 1799. Proust discovered this law while analyzing samples of Cupric Carbonate. This law states that all pure samples of a particular chemical compound contain similar elements combined in the same proportion by mass. This law is based on the fact that when elements combine to form a given compound, they do so in fixed proportions by mass so that all pure samples of that compound are identical in composition

Experiment to Verify the Law

Prepare pure samples of cupric oxide by two different methods

1. By heating copper carbonate
2. By the decomposition of cupric nitrate

The cupric oxide prepared by both the methods always contains the same elements copper and oxygen combined together in the same fixed proportion of 4: 1 by weight. This illustrates the law of definite proportions. It can be verified by taking a known weight of a pure sample (W_1 gm) of cupric oxide in a porcelain boat.

$2\ Cu(NO_3)_2 \rightarrow 2\ CuO + 4\ NO_2 \uparrow + O_2 \uparrow$

It is placed inside a hard glass tube kept horizontally. A current of pure dry hydrogen is sent inside the tube and the tube is heated. The cupric oxide is reduced to metallic copper.

$CuO + H_2 \rightarrow Cu + H_2O$

The Weight of copper formed is found out W_2 gm.

Calculation:

Method 1:

Weight of cupric oxide = W_1 gm

Weight of Copper = W_2 gm

∴ Weight of oxygen = $W_1 - W_2$ gm

Ratio of copper: oxygen = $W_2 : (W_1 - W_2)$

The same experiment is repeated with a known weight W_3 gm of cupric oxide prepared by heating copper carbonate

CuCO3 ——> $CuO + CO_2\uparrow$

The cupric oxide formed is reduced to metallic copper by passing a current of pure and dry hydrogen inside the tube as before. The weight of metallic copper was found to be W_4 gm. The ratio of the weight of copper to the weight of oxygen in both the samples is calculated as follows:

Method 2:

Weight of cupric oxide = W_3 gm.

Weight of copper = W_4 gm

∴ Weight of oxygen = $W_3 - W_4$ gm

Ratio of copper to oxygen = $W_4 : (W_3 - W_4)$

The two ratios $[W_2 : W_1 - W_2]$ and $[W_4 : W_3 - W_4]$ are found to be the same and is equal to 4: 1. Thus the law of definite proportions is verified experimentally

Example 1: 1.375 g of CuO were reduced by H and 1.098 g of Cu were obtained. In another experiment, 1.178 g of Cu were dissolved in nitric acid and the resulting copper nitrate was converted into CuO by ignition. The weight of CuO formed was 1.476 g. Show that these results prove the law of constant proportion.

Solution

Experiment 1:

Weight of CuO = 1.375 g

Weight of Cu = 1.098 g

Weight of oxygen = (1.375 – 1.098) g = 0.277 g

Ratio of copper oxygen = 1.098 : 0.277 = 3.96 : 1

Experiment 2:

Weight of CuO = 1.476

Weight of Cu = 1.178

Weight of oxygen = 1.476 – 1.178 = 0.298

Ratio of copper: oxygen = 1.178 : 0.298 = 3.96 : 1

In both experiments the ratio of Copper: oxygen is some (3.96: 1). Hence it illustrates the law of definite proportions.

Example 2: In an experiment 0.2430 gm of magnesium on burning with oxygen yielded 0.4030 gm of magnesium oxide. In another experiment 0.1820 gm of magnesium on burning with oxygen yielded 0.3020 gm of magnesium oxide. Show that the data explain the law of definite proportions.

Solution

Experiment 1:

Weight of Magnesium oxide = 0.4030 gm

Weight of Magnesium = 0.2430 gm

Weight of oxygen = 0.4030 – 0.2430 = 0.16 gm

Ratio of Magnesium: oxygen = 0.2430: 0.16 = 1.552: 1

Experiment 2:

Weight of Magnesium oxide = 0.3020

Weight of Magnesium = 0.1820

Weight of oxygen = 0.3020 – 0.1820 = 0.12

Ratio of magnesium: oxygen= 0.1820: 0.12 = 1.552: 1

In both experiments the ratio of magnesium: oxygen is same (1.518:1) Hence it illustrates the law of definite proportions.

The Law of Definite Proportion or Constant Composition states that all pure samples of the same chemical compound contain the same chemical compound contain the same elements combined in the same proportion by mass.

Law of Multiple Proportions
This Law of Multiple Proportions was enunciated by John Dalton at about the same time he postulated his Atomic Theory of Matter in 1803. This law states that when two elements A and B combine to form two or more compounds, then different weights of B which combine with a fixed weight of A bears a simple numerical ratio to one another. The law states that if two elements A and B, combine to form more than one chemical compound, then the various masses of one element A, which combine separately with a fixed mass of the other element B are in simple multiple proportion.

Carbon combines with oxygen to form two different oxides, namely, carbon monoxide (CO) and carbon dioxide (CO_2). The proportions by weight of the two elements are:

Carbon monoxide – C: O :: 12 : 16

Carbon dioxide – C: O :: 12 : 32

There, the weights of oxygen that combine with a fixed weight of carbon (12g) are in the ratio 16g : 32g i.e. 1:2, a simple numerical ratio.

Experiment to Verify the Law

The law can easily be verified by the study of oxides of copper. Copper reacts with oxygen to form two oxides – the red cuprous oxide (Cu2O) and the black cupric oxide (CuO). In order to verify the law of multiple proportions, fixed amounts of these oxides (say 20g each) are separately reduced to metallic copper by heating them in a current of hydrogen and the masses of copper obtained from them are estimated. The difference in the mass of oxide taken and the mass of copper obtained from it gives the mass of oxygen present in it.

Now the masses of oxygen which combine with a definite mass of copper in the two oxides are calculated. These masses are found in a simple whole number ratio. This verifies the law of multiple proportions.

Note: The law is valid till an element is present in one particular isotopic form in all its compounds. When an element exists in the form of different isotopes in its compounds, the law does not hold good.

Example: In an experiment, 34.5 g oxide of a metal was heated so that O_2 was liberated and 32.1 g of metal was obtained. In another experiment 119.5 g of another oxide of the same metal was heated and 103.9 g metal was obtained and O_2 was liberated. Calculate the mass of O_2 liberated in each experiment. Show that the data explain the law of multiple proportions.

Solution

Experiment 1

Weight of the metal oxide = 34.5 g

Weight of the metal = 32.1 g

32.1 g metal combines with 2.4 g oxygen

1 g of the metal combine with 2.4/32.1 = 0.075 g

Experiment 2

Weight of the oxide taken = 119.5 g

Weight of the metal formed = 103.9 g

Weight of oxygen liberated = 15.6 g

103.9 g of metal combines with 55.6 g oxygen.

1 g of metal = ——– 15.6/103 x 1 = 0.15014 oxygen

Therefore different weights of oxygen that combine with the fixed weight of the metal (1 g) are in the ratio:

0.1501 : 0.075

2 : 1

The proportion by weight of oxygen is indicated by simple ratio. Thus law of multiple proportions is obeyed.

Law of Reciprocal Proportions

The Law of Reciprocal proportion states that the masses of several elements A, B, and C which combine separately with a fixed mass to form D are the same as or simple multiples of the masses in which A, B, C, themselves combine with one another.

Assessment

1. How many atoms are contained in 1 mole of hydrogen molecule.
   a. 18.09×10^{23} atoms
   b. 12.06×10^{23} atoms
   c. 6.02×10^{23} atoms
   d. 6.02×10^{23} molecules

2. The percentage of Oxygen in Sulphur IV Oxide (S = 32, O = 16) is
   a. 5%
   b. 50%
   c. 500%
   d. 25%

3. If the relative molecular mass of CH_2O is 60, calculate its empirical formula. (C = 12, H = 1, O = 16)
   a. 4
   b. 1
   c. 2
   d. 3

4. Find the empirical formulae of the following compounds from their percentage composition by mass
   a. N = 26.17%, H = 7.48%, Cl = 66.35%
   b. Ca = 71.43%, O = 28.57%
   c. Ag = 63.53%, N = 8.23%, O = 28.24%
   d. Na = 32.40%, O = 45.07%, S = 22.53%
   (N = 14, H = 1, Cl = 35.5, Ca = 40, O = 16, Ag = 108, Na = 23, S = 32)

Answers

1. B
2. B
3. C

Week 6

Topic: Chemical Combinations or Bonding

Introduction

A chemical bond is the result of an attraction between atoms or ions. Chemical compounds are formed by the joining of two or more atoms. The types of bonds that a molecule contains will determine its physical properties, such as melting point, hardness, electrical and thermal conductivity, and solubility.

The two extreme cases of chemical bonds are:

- Covalent bond: bond in which one or more pairs of electrons are shared by two atoms.
- Ionic bond: bond in which one or more electrons from one atom are removed and attached to another atom, resulting in positive and negative ions which attract each other.

Other types of bonds include co-ordinate covalent, metallic bonds and hydrogen bonding. The attractive forces between molecules in a liquid can be characterized as Van der Waals bonds.

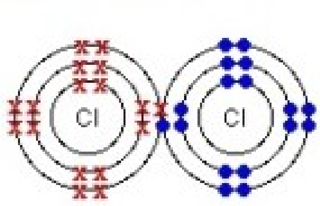
Covalent Bonds

Covalent chemical bonds involve the sharing of a pair of valence electrons by two atoms, in contrast to the transfer of electrons in ionic bonds. Such bonds lead to stable molecules if they share electrons in such a way as to create a noble gas configuration for each atom.

Hydrogen gas forms the simplest covalent bond in the diatomic hydrogen molecule. The halogens such as chlorine also exist as diatomic gases by forming covalent bonds. The nitrogen and oxygen which makes up the bulk of the atmosphere also exhibits covalent bonding in forming diatomic molecules.

Examples of covalent bonding

Chlorine

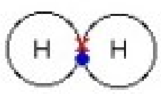


For example, two chlorine atoms could both achieve stable structures by sharing their single unpaired electron as in the diagram.

The two chlorine atoms are said to be joined by a covalent bond. The reason that the two chlorine atoms stick together is that the shared pair of electrons is attracted to the nucleus of both chlorine atoms.

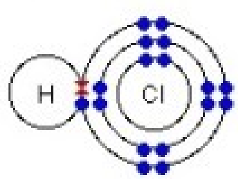
Note: The fact that one chlorine has been drawn with electrons marked as crosses and the other as dots is simply to show where all the electrons come from. In reality there is no difference between them.

Hydrogen



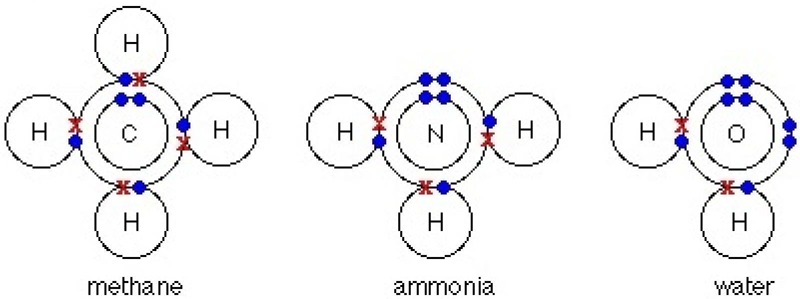
Hydrogen atoms only need two electrons in their outer level to reach the noble gas structure of helium. Once again, the covalent bond holds the two atoms together because the pair of electrons is attracted to both nuclei.

Hydrogen chloride

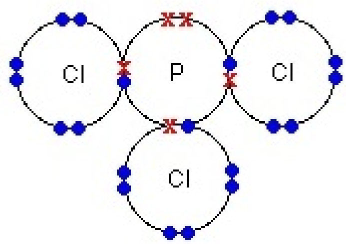


The hydrogen has a helium structure, and the chlorine an argon structure

Other examples:



Even with a more complicated molecule like PCl_3, there's no problem. In this case, only the outer electrons are shown for simplicity. Each atom in this structure has inner layers of electrons of 2,8. Again, everything present has a noble gas structure.

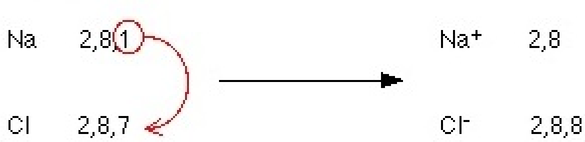


Ionic Bonds (Electrovalent Bonds)

In an ionic bond, an electron is actually transferred from the less electronegative atom (metals) to the more electronegative element (non-metals). Ionic bonds are the result of an electrostatic attraction between ions that have opposite charges; in other words, cations and anions. Ionic bonds usually form between metals and nonmetals. Ionic bonds are very strong, so compounds that contain these types of bonds have high melting points and exist in a solid state under standard conditions.

Ionic bonding in sodium chloride

Sodium (2,8,1) has 1 electron more than a stable noble gas structure (2,8). If it gave away that electron it would become more stable.



Chlorine (2,8,7) has 1 electron short of a stable noble gas structure (2,8,8). If it could gain an electron from somewhere it too would become more stable.

The answer is obvious. If a sodium atom gives an electron to a chlorine atom, both become more stable.

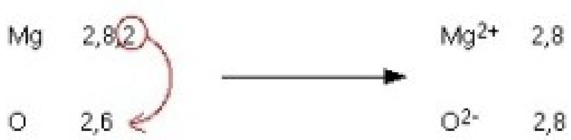
The sodium has lost an electron, so it no longer has equal numbers of electrons and protons. Because it has one more proton than electron, it has a charge of 1+. If electrons are lost from an atom, positive ions are formed. Positive ions are sometimes called cations.

The chlorine has gained an electron, so it now has one more electron than proton. It therefore has a charge of 1-. If electrons are gained by an atom, negative ions are formed. A negative ion is sometimes called an anion.

The sodium ions and chloride ions are held together by the strong electrostatic attractions between the positive and negative charges. One sodium atom to provide the extra electron for one chlorine atom, so they combine together 1:1. The formula is therefore NaCl.

Some other examples of ionic bonding

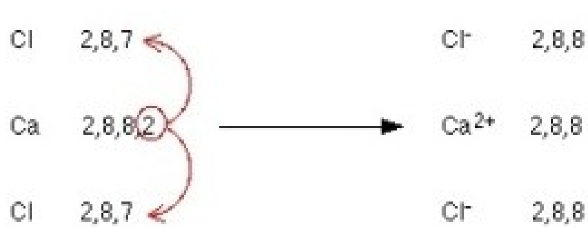
Magnesium Oxide



Again, noble gas structures are formed, and the magnesium oxide is held together by very strong attractions between the ions. The ionic bonding is stronger than in sodium chloride because this time you have 2+ ions attracting 2- ions. The greater the charge, the greater the attraction.

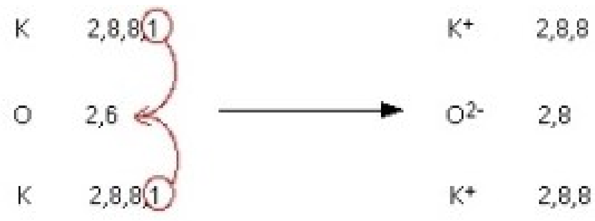
The formula of magnesium oxide is MgO.

Calcium Chloride



This time you need two chlorines to use up the two outer electrons in the calcium. The formula of calcium chloride is therefore $CaCl_2$.

Potassium Oxide



Again, noble gas structures are formed. It takes two potassium to supply the electrons the oxygen needs. The formula of potassium oxide is K_2O.

Co-ordinate (dative covalent) bonding

A covalent bond is formed by two atoms sharing a pair of electrons. The atoms are held together because the electron pair is attracted by both of the nuclei.

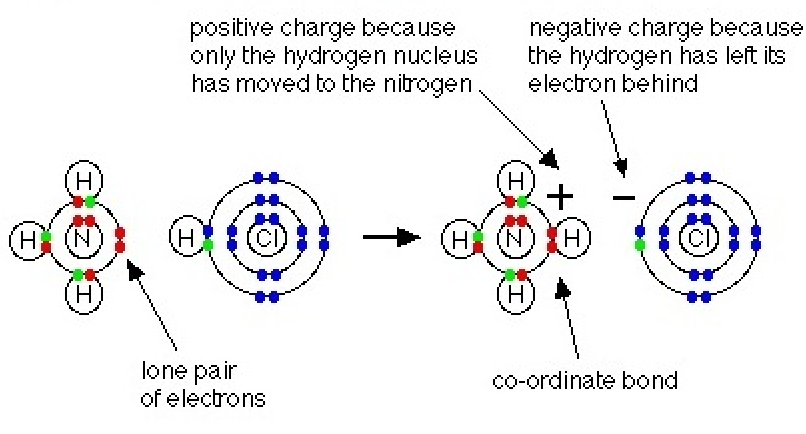
In the formation of a simple covalent bond, each atom supplies one electron to the bond – but that doesn't have to be the case. A co-ordinate bond (also called a dative covalent bond) is a covalent bond (a shared pair of electrons) in which both electrons come from the same atom.

The reaction between ammonia and hydrogen chloride

If these colourless gases are allowed to mix, a thick white smoke of solid ammonium chloride is formed.

$$NH_3 + HCl \longrightarrow NH_4Cl$$

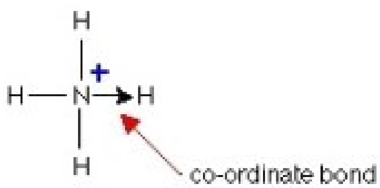
Ammonium ions, NH_4^+, are formed by the transfer of a hydrogen ion from the hydrogen chloride to the lone pair of electrons on the ammonia molecule.



When the ammonium ion, NH_4^+, is formed, the fourth hydrogen is attached by a dative covalent bond, because only the hydrogen's nucleus is transferred from the chlorine to the nitrogen. The hydrogen's electron is left behind on the chlorine to form a negative chloride ion.

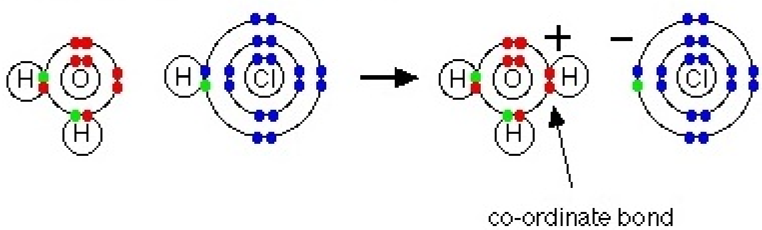
Once the ammonium ion has been formed it is impossible to tell any difference between the dative covalent and the ordinary covalent bonds.

In simple diagrams, a co-ordinate bond is shown by an arrow. The arrow points from the atom donating the lone pair to the atom accepting it.



Another example is dissolving hydrogen chloride in water to make hydrochloric acid

Something similar happens. A hydrogen ion (H^+) is transferred from the chlorine to one of the lone pairs on the oxygen atom.



$$H_2O + HCl \longrightarrow H_3O^+ + Cl^-$$

Metallic Bonds

Metallic bond is the bond that holds atom together in a metal. Metallic bond occurs as a result of the attraction between delocalised electrons and the positive metal atom nuclei. The properties of metals suggest that their atoms possess strong bonds, yet the ease of conduction of heat and electricity suggest that electrons can move freely in all directions in a metal.

Metal Properties

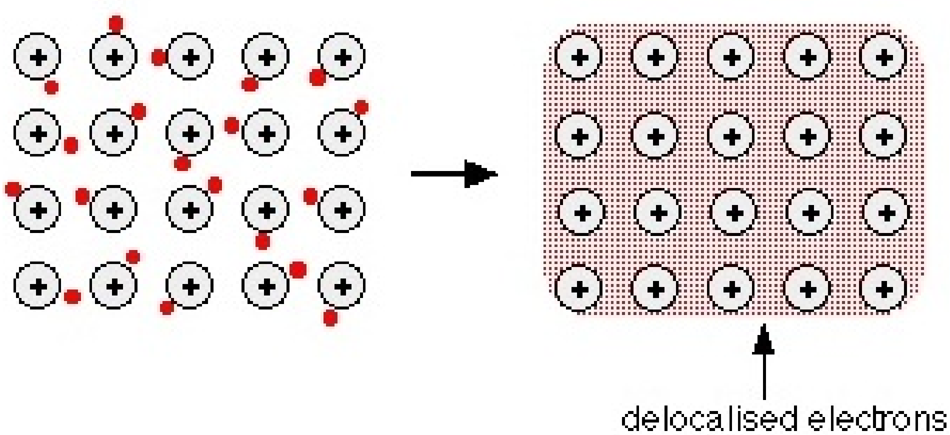
The general properties of metals include malleability and ductility and most are strong and durable. They are good conductors of heat and electricity. Their strength indicates that the atoms are difficult to separate, but malleability and ductility suggest that the atoms are relatively easy to move in various directions. The electrical conductivity suggests that it is easy to move electrons in any direction in these materials. The thermal conductivity also involves the motion of electrons. All of these properties suggest the nature of the metallic bonds between atoms.

Metallic bonding in sodium

Metals tend to have high melting points and boiling points suggesting strong bonds between the atoms. Even a metal like sodium (melting point 97.8°C) melts at a considerably higher temperature than the element (neon) which precedes it in the Periodic Table.

Sodium has the electronic structure $1s^2 2s^2 2p^6 3s^1$. When sodium atoms come together, the electron in the 3s atomic orbital of one sodium atom shares space with the corresponding electron on a neighbouring atom to form a molecular orbital – in much the same sort of way that a covalent bond is formed.

The electrons can move freely within these molecular orbitals, and so each electron becomes detached from its parent atom. The electrons are said to be delocalised. The metal is held together by the strong forces of attraction between the positive nuclei and the delocalised electrons.



This is sometimes described as "an array of positive ions in a sea of electrons".

Each positive centre in the diagram represents all the rest of the atom apart from the outer electron, but that electron hasn't been lost – 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^+.

Metallic bonding in magnesium

Magnesium has the outer electronic structure $3s^2$. Both of these electrons become delocalised, so the "sea" has twice the electron density as it does in sodium. Each magnesium atom has 12 protons in the nucleus compared with sodium's 11. In both cases, the nucleus is screened from the delocalised electrons by the same number of inner electrons – the 10 electrons in the $1s^2\ 2s^2\ 2p^6$ orbitals. That means that there will be a net pull from the magnesium nucleus of 2+, but only 1+ from the sodium nucleus. So not only will there be a greater number of delocalised electrons in magnesium, but there will also be a greater attraction for them from the magnesium nuclei.

Magnesium atoms also have a slightly smaller radius than sodium atoms, and so the delocalised electrons are closer to the nuclei.

Metallic bonding in transition elements

Transition metals tend to have particularly high melting points and boiling points. The reason is that they can involve the 3d electrons in the delocalisation as well as the 4s. The more electrons you can involve, the stronger the attractions tend to be.

The metallic bond in molten metals

In a molten metal, the metallic bond is still present, although the ordered structure has been broken down. The metallic bond isn't fully broken until the metal boils. That means that boiling point is actually a better guide to the strength of the metallic bond than melting point is. On melting, the bond is loosened, not broken.

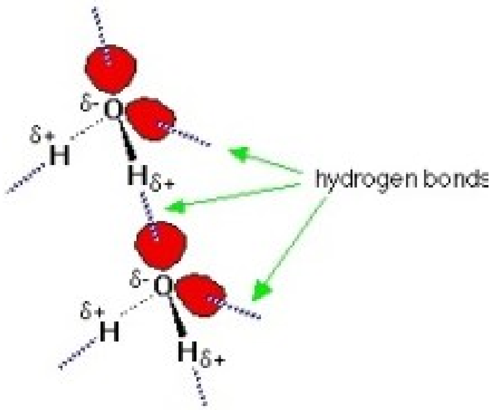
Weak Bonds

Hydrogen Bonding

This is an intermolecular force which arises when hydrogen is covalently linked to elements like nitrogen, oxygen and fluorine. These latter elements are strongly electronegative i.e they have very strong affinity for electrons. They require only a few electrons to attain their octet configuration. They tend to pull the shared electrons of covalent bonds towards themselves which results in the formation of dipole where the hydrogen atom is partially positive while the nitrogen, oxygen or fluorine is partially negative. This is a dipole-dipole attraction which occurs between a hydrogen atom attached to a strongly electronegative atom. Hydrogen bonding differs from other uses of the word "bond" since it is a force of attraction between a hydrogen atom in one molecule and a small atom of high electronegativity in another molecule. That is, it is an intermolecular force, not an intramolecular force as in the common use of the word bond.

If the hydrogen is close to another atom such as oxygen, fluorine or nitrogen in another molecule, then there is a force of attraction termed a dipole-dipole interaction. This attraction or "hydrogen bond" can have about 5% to 10% of the strength of a covalent bond.

Consider two water molecules coming close together.

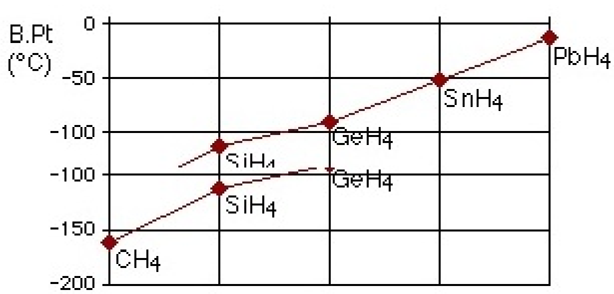


The + hydrogen is so strongly attracted to the lone pair that it is almost as if you were beginning to form a co-ordinate (dative covalent) bond.

Hydrogen bonding has a very important effect on the properties of water and ice. Hydrogen bonding is also very important in proteins and nucleic acids and therefore in life processes. The “unzipping” of DNA is a breaking of hydrogen bonds which help hold the two strands of the double helix together.

The evidence for hydrogen bonding

Many elements form compounds with hydrogen. If you plot the boiling points of the compounds of the Group 4 elements with hydrogen, you find that the boiling points increase as you go down the group.



The increase in boiling point happens because the molecules are getting larger with more electrons, and so Vander Waals dispersion forces become greater.

Vander Waal’s Force

This is the type of force that arises in gases such as nitrogen, oxygen, chlorine which are diatomic and non-polar molecules. They are intermolecular attraction and not real bond.

Vander Waal's force increases with increase in the number of electrons. They are stronger in iodine (solid at room temperature) than in bromine (liquid at room temperature) and less in chlorine (gas at room temperature). Weak attractive forces exist even between discrete molecules. These forces are very weak when compared with ionic and covalent bonds but they are important in liquefaction of gases and in formation of molecular lattices in iodine and naphthalene crystals.

Assessment

1. The major reason a chemical reaction occurs among elements is that they have the tendency to
   a. attain the nearest noble gas structure
   b. become a metal
   c. become a non-metal
   d. become any noble element

2. In electrovalency most metallic atoms with few valence electrons give out electrons because
   a. they are unstable
   b. they require less energy to give away these electrons
   c. they require more energy to give away these electrons
   d. they need non-metals to operate

3. In electrovalency, valence electrons are transferred and the atomic number is
   a. also reduced
   b. stabilized
   c. unaffected
   d. destabilized

4. The bond type in a diatomic Nitrogen gas is
   a. double covalent
   b triple covalent
   c. single covalent
   d. double electrovalent bond

5. The bond between two iodine molecules are
   a. co-ordinate bond
   b. electrovalent bond
   c. ionic bond
   d. van der Waal's bond

6. Bonds between a highly electronegative atom and a hydrogen from another molecule is called

a. hydrogen bond
b. covalent bond
c. inter molecular forces
d. ligand

Answers

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

Week 7

Topic: The Kinetic Theory of Matter

Introduction

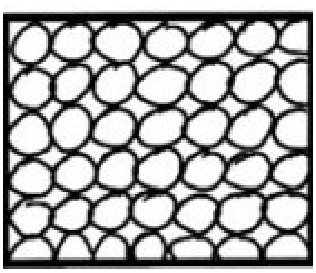
The Kinetic Theory of Matter states that matter is composed of a large number of small particles that are in constant motion. The law states that the tiny particles of matter are continually moving and so possess kinetic energy. In other words, kinetic theory of matter recognizes that matter is composed of very small particles (ions, atoms and molecules) whose different pattern of arrangements and motions result in the different possible states in which matter can occur. It also explains the properties of these states. An increase in temperature causes an increase in average kinetic energy . This theory is also called the Kinetic Molecular Theory of Matter and the Kinetic Theory.

There are three states of matter; these are solid, liquid and gas.

Substances can change from one state to another. Kinetic theory can explain the change of state by considering all matter (substances) to be made of particles.

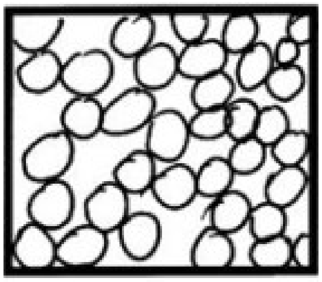
Solid State

In a solid, particles are closely packed in a regular arrangement and are unable to move about. The particles are held together by cohesive forces. The cohesive forces may be electrovalent, covalent or even Vander Waal's force. The particles vibrate about a fixed position and have a definite shape and definite volumes.



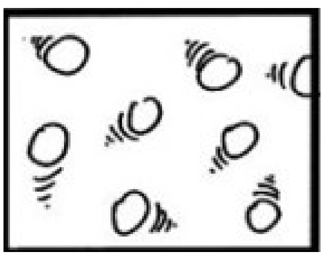
Liquid State

The particles are closely packed in a random arrangement. The particles can move through the liquid because they have more kinetic energy than solid and they are no longer held in a fixed position but they cling together. A liquid does not have a fixed shape but normally takes the shape of its container.



Gas State

The particles are far apart. There motion is random and independent of the other particles. The forces of attraction between particles in a gas are very weak, so the particles are free to move about in all directions at great speed. They can be restricted only by the walls of the container. Because of the large spaces between the molecules, a gas can be compressed easily.



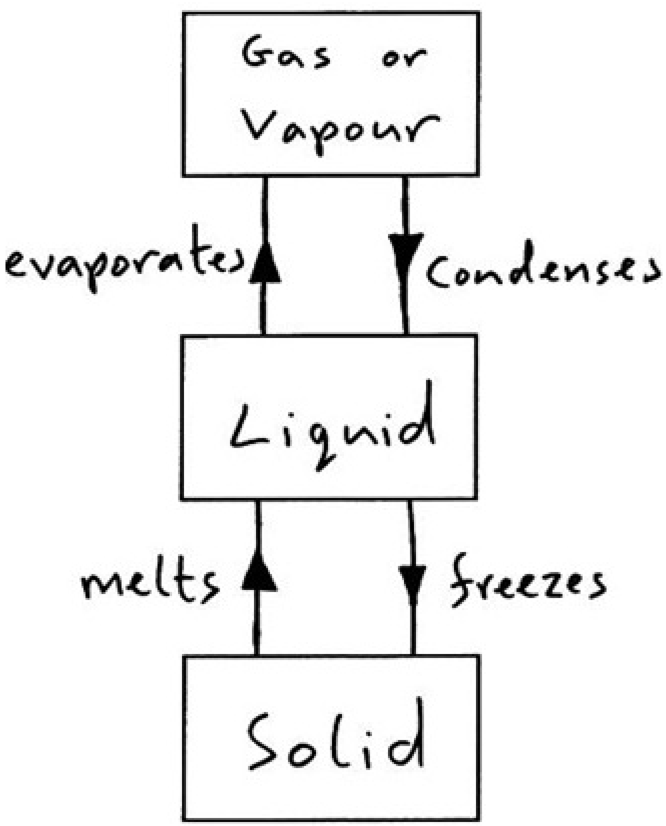
The main aspects of the kinetic theory are:

- Matter is composed of very tiny particles (atoms or molecules), which are separated from each other by inter-particle distances.
- Each particle of matter is in constant motion.
- In a gas, the particles can move around freely and independently.
- In a liquid, particle movement is a bit constrained and limited to sliding/flow movement within its volume.
- In a solid, particle movement is fully constrained and restricted to only vibrational motion of particles in their fixed positions within the solid.
- The particles of matter experience forces of attraction amongst themselves. These attractive forces decrease rapidly with increasing distance between the particles.
- Particles in solids are very close to each other, and the attractive forces are large enough to hold the particles in fixed positions. Thus, a solid has a fixed shape and a fixed size (volume).

- The particles of liquids are a little further apart and are free to slide and flow, taking the shape of the container. Thus, a liquid has no fixed shape. However, since the particle movement is restricted to within the space occupied by the liquid, a liquid does have a fixed size (volume).
- The separations between particles of a gas are quite large, resulting in complete freedom of motion. Hence, a gas has neither fixed shape nor fixed size (volume), and tends to expand to fill up the entire volume of its container.
- Because the particles are in motion, they possess kinetic energy. The temperature of matter is a measure of the average kinetic energy possessed by the particles. When heat is applied to matter, it gets absorbed and translated to increased kinetic energy of the particles (which means greater motion), resulting in a rise in temperature.

Change of State and Kinetic Theory

The kinetic theory of matter gives a clear explanation of the internal processes involved at the particle level when matter undergoes a change of state.



A given substance can exist as a solid, liquid or gas. Change of state is brought about by a change in temperature. When a substance is heated, its particles acquire more kinetic energy and when cooled they become less energetic.

Process of Heating

Theoretically, heating a solid to higher and higher temperatures changes its phase to a liquid, and finally to a gas.

- Fusion (solid to liquid)

A solid consists of low kinetic energy vibrating particles locked into position by interparticle attractive forces. When heat is applied, energy is absorbed and the particles start vibrating more vigorously. Finally, the vibrations become energetic enough to overcome the attractive forces, and the particles start sliding out of their positions to flow about. The solid is now melting into a liquid.

- Vaporisation (liquid to gas)

On further application of heat to the liquid, the particles move around more energetically within the volume of the liquid. Finally, they become energetic enough to start escaping from the surface of the liquid, overcoming the backward pull by their neighbours in the volume of the liquid. The process of boiling has begun, wherein the liquid converts to gas as particles escape to move around independently without any constraints.

- Evaporation (liquid to gas)

According to the kinetic theory, the temperature is a measure of the average kinetic energy possessed by the particles of matter. This means that in any sample of matter, there will be particles with higher kinetic energy than average, balanced by those with lower energy than average. So, even in a liquid whose temperature is not high enough for boiling to occur, there will be some particles with sufficient kinetic energy to break through the surface of the liquid overcoming the backward pull of others. They slowly escape as gas particles, and the process is called evaporation.

Process of Cooling

Generally, cooling a gas changes its phase to a liquid, and finally to a solid.

- Condensation (gas to liquid)

When a gas is cooled (i.e. heat is removed) progressively, the free moving particles start losing kinetic energy and slowing down. Finally, the forces of attraction between the lower energy particles colliding with each other are strong enough to hold them together, and the gas begins to condense into liquid.

- Solidification (liquid to solid)

The particles still have energy enough to slide about within the volume of the liquid, but further cooling lowers this energy further.

Application of the Kinetic Theory of Matter to Explain the Nature of Gases

The three properties of gases that are especially important are diffusibility, thermal expansion and compressibility. All gases are characterized by diffusibility, but the rates at which different gases diffuse depend on their molecular weights.

When heated, gases expand to a much greater extent than do solids or liquids- all gases tend to behave alike in this respect. In comparison with solids and liquids, gases are very easily compressed – all gases tend to behave alike in this regard also. These properties, as well as the empirical laws governing the behaviour of gases can be explained by the kinetic theory.

- Explanation of Diffusion of Gases by the Kinetic Theory

Diffusion is a phenomenon whereby particles of a substance move from an area of high concentration into an area of low concentration. Gases diffuse rapidly. For example, if a small quantity of an odorous gas, e.g. hydrogen sulphide, is released at one point of a room, the smell soon gets to all parts of the room. This can be explained using the kinetic theory of gases.

From the assumptions of the theory, we have that:

Gases are made of discrete particles called molecules, and not a single piece. If they were made of a single piece, then, the smell of the hydrogen sulphide would not pervade the whole room at the same time, but would probably be perceived at one corner of the room at a time.

The molecules are relatively far apart and are in rapid, random motion, moving at high speeds in straight lines. This account for the smell getting to every part of the room in a couple of minutes after the release.

- Explanation of Compressibility of Gases By the Kinetic Theory

Compressibility of gases can be explained from the assumption of the kinetic theory, which states that a gas consists of particles that are separated from one another by large spaces. Based on this, it is therefore easy to bring the molecules closer together (i.e. compressed) when the volume of the container is reduced.

Reduction in volume leads to decrease in temperature (according to Charles' law, $V \alpha T$). Hence, compression of gases results in a drop of temperature in the system – the kinetic energy of the system also drops.

- Explanation of Expansion of Gases by the Kinetic Theory

Expansion of gases can be explained by the kinetic theory from the assumptions which state that: Gases are in constant rapid motion, moving at great speeds, occupying the volume of the container.

The average kinetic energy of all the molecules is assumed to be directly proportional to the absolute temperature of the gas. The greater the average kinetic energy of gas molecules, the greater they are able to move and occupy more volume. Therefore, at higher temperatures, gases obtain higher kinetic energy, and thus expand (or occupy large volumes).

Assumptions of the Kinetic Theory of Gases

This theory describes the behaviour of an Ideal gas.

1. The gas molecules move randomly in straight lines colliding with each other and with the walls of the container. The collisions of the gas molecules on the walls of the container constitute the gas pressure exerted on the container

2. The collisions of the gas molecules are perfectly elastic. When two individuals collide, their individual energies may change and one may move faster than the other but the total kinetic energy remains the same

3. The actual volume occupied by the gas molecules themselves are negligible relative to the volume of the container

4. The cohesive forces between the gas molecules are negligible

5. The temperature of the gas is a measure of the average kinetic energy of the gas molecules.

Assessment

1. Water exists as a solid, liquid and gas respectively because
   a. water is colourless
   b. water is electrovalent
   c. water in any state possesses a certain degree of motion in molecules
   d. water is molecular

2. Which of the three states of matter has no fixed shape, no fixed volume and less dense
   a. Gas
   b. Liquid
   c. Solid
   d. Crystal

3. Presence of Sodium Chloride in ice will
   a. decrease or lower the boiling point of sodium chloride
   b. increase the melting point of sodium chloride
   c. make sodium chloride impure
   d. lower the freezing point of sodium chloride

4. The phenomenon whereby the atmospheric pressure equals the saturated vapour pressure is called
   a. freezing
   b. latent heat
   c. boiling
   d. normal pressure

5. The escape of molecules with more than average kinetic energy of the molecules is called
   a. melting
   b. freezing
   c. evaporation
   d. efflorescence

Answers

1. C
2. A
3. A
4. C
5. C

Week 8

Topic: Gas Laws

Introduction

Gas has existed since the beginning of time; oftentimes, it was referred to as “air” or “oxygen;” however, in the late 18th century, “air” became known as gas, and people were able to distinguish between different types of gas. Towards the end of the 18th century, scientists started testing and developing laws that later became known as the “gas laws.” One of the most amazing things about gases is that, despite wide differences in chemical properties, all the gases more or less obey the gas laws. These laws describe properties of gases, that is how gases behave with respect to pressure, volume, temperature and how they react in different situations. In order to understand the gas laws, we need to define a few terms:

Gas: A substance consisting of widely spread particles; it can expand indefinitely. This is also the third state of matter; it is not a solid or a liquid.

Pressure: The measure of force applied by another substance (such as a gas). It is commonly abbreviated as “P” (a capital letter P). Pressure can be measured in millimeters of Mercury (mmHg), torr, atmospheres (atm), Pascals (Pa), and kilopascals (kPa). All of the following measurements are the same, just different units, so you can use them to convert from one to the other. For the ideal gas law, the pressure will need to be in atmospheres. The conversions between these are as follows:

760 mmHg = 760 torr = 1.00 atm = 101,325 Pa = 101.325 kPa

Volume: The numerical amount of space occupied by a solid, liquid, or gas. It is commonly abbreviated as “V” (a capital letter V). Volume, in this situation, will be most often measured in liters, L.

Temperature: The measurement of the amount of energy seen in the motion of particles in a solid, liquid or gas. It can be measured on three scales: Fahrenheit, Celsius (sometimes referred to as Centigrade) and Kelvin. It is commonly abbreviated as “T” (a capital letter T). Temperature, in this situation, will most often be measured in Kelvin, K.

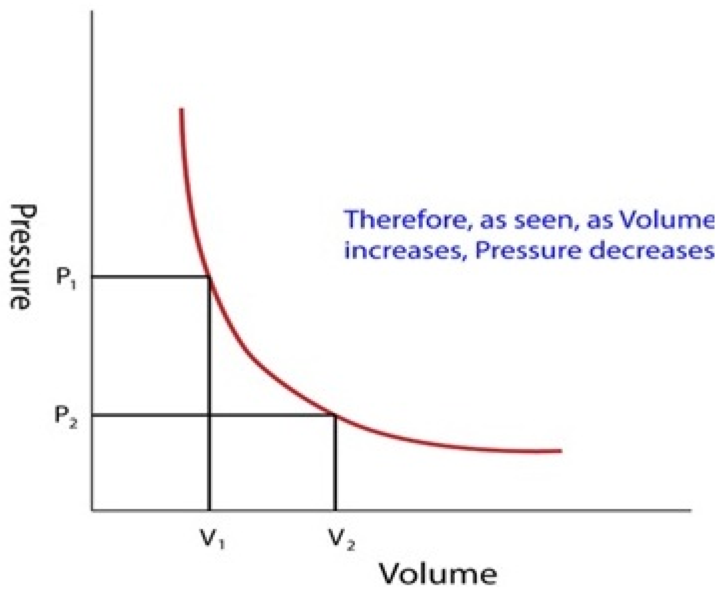
STP: STP stands for “standard temperature and pressure” and refers to conditions of 273 K (0 degrees C) and 1 atm.

The Gas Laws: Pressure Volume Temperature Relationships

Boyle’s Law: The Pressure-Volume Law

Boyle's law or the pressure-volume law states that the volume of a given amount of gas held at constant temperature varies inversely with the applied pressure when the temperature and mass are constant.

$$V \propto \frac{1}{P}$$



PV = C (C = Constant)

When pressure goes up, volume goes down. When volume goes up, pressure goes down. From the equation above, this can be derived:

$P_1V_1 = P_2V_2$

Or like this:

Example 1: If the initial volume was 500 mL at a pressure of 760 atm, when the volume is compressed to 450 ml, what is the pressure?
Solution: V_1 = 500mL, P_1 = 760atm, V_2 = 450mL, P_2 = ?

$P_1V_1 = P_2V_2$

$(760)(500) = P_2(450)$
$760 \times 500 / 450 = P_2$

P_2 = 844 atm
The pressure is 844 atm after compression.

Example 2: A 17.50cm^3 sample of gas is at 4.50 atm. What will be the volume if the pressure becomes 1.50 atm, with a fixed amount of gas and temperature?

Solution: V_1 = 17.50cm3, P_1 = 4.50atm, V_2 = ?, P_2 = 1.50atm

$P_1V_1 = P_2V_2$

$(4.50)(17.50) = (1.50)V_2$

$4.50 \times 17.50 / 1.50 = V_2$

$V_2 = 52.50cm^3$

Example 3: A sample of air occupies a volume of 450 L at 20 °C and 100 mmHg. What will be the pressure of this gas if it is transferred to a 200 L bulb at the same temperature?

Solution: V_1 = 450 L, P_1 = 100 mmHg, P_2 = ?, V_2 = 200 L

$(100)(450) = P_2(200)$

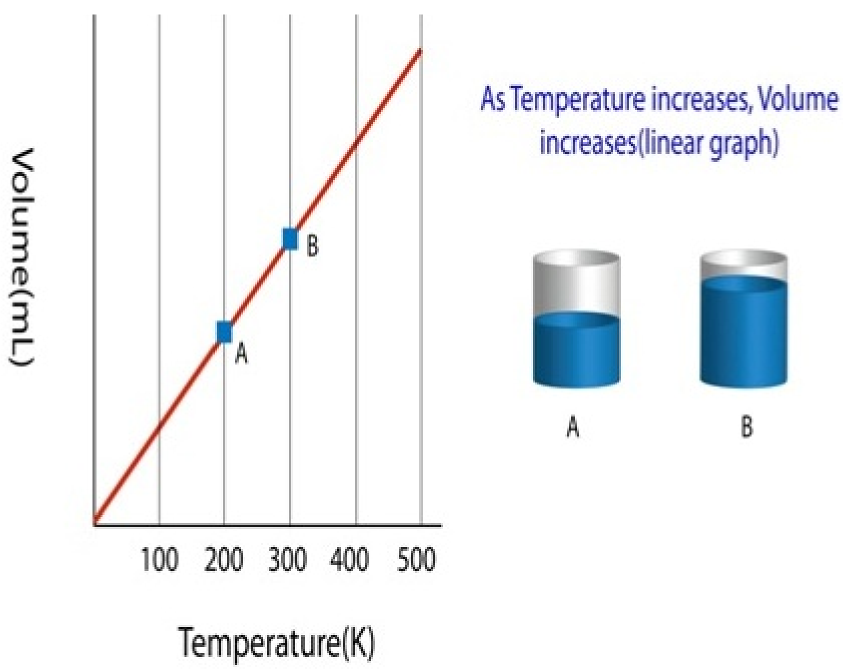
$100 \times 450 / 200 = P_2$

P_2 = 225 L

Charles' Law: The Temperature-Volume Law

This law states that the volume of a given amount of gas held at constant pressure is directly proportional to the temperature (Kelvin).

V T



Same as before, a constant can be put in:

$V / T = C$ (C – Constant)

As the volume goes up, the temperature also goes up, and vice-versa.
Also same as before, initial and final volumes and temperatures under constant pressure can be calculated.

$V_1 / T_1 = V_2 / T_2$

Or like this:

Example 1: A sample of Carbon dioxide in a pump has volume of 20.5 cm^3 and it is at 40 °C. When the amount of gas and pressure remain constant, find the new volume of Carbon dioxide in the pump if temperature is increased to 65 °C.

Solution: V_1 = 20.5 cm^3, T_1 = 40 °C = (273 + 40)K, V_2 = ?, T_2 = 65 °C = (273 + 65)K

$V_1 / T_1 = V_2 / T_2$

$V_2 = V_1 . T_2 / T_1$

$V_2 = 20.5 \times 338 / 313$

$= 22.1\ cm^3$

Example 2: A sample of gas occupies 400.0 mL at 25.00 °C and 1 bar pressure. What volume will it occupy at 200.00 °C at the same P?

Solution: V_1 = 400.0 mL, T_1 = 25.00 °C = (25.00 + 273), V_2 = ?, T_2 = 200.00 °C = (200.00 + 273)K

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow V_2 = V_1 \frac{T_2}{T_1}$$

$$V_2 = 400.0\ mL \frac{(200.00 + 273.15)K}{(25.00 + 273.15)K} = 634.8\ mL$$

Gay-Lussac's Law: The Pressure Temperature Law

This law states that the pressure of a given amount of gas held at constant volume is directly proportional to the Kelvin temperature.

P T

Same as before, a constant can be put in:

$P / T = C$

As the pressure goes up, the temperature also goes up, and vice-versa.
Also same as before, initial and final volumes and temperatures under constant pressure can be calculated.

$P_1 / T_1 = P_2 / T_2$

Example 1: Find the final pressure of gas at 150 K, if the pressure of gas is 210 mmHg at 120 K if the volume remains constant.

Solution: T_2 = 150K, P_2 = ?, P_1 = 210 mmHg, T_1 = 120K

$P_1 / T_1 = P_2 / T_2$

$P_2 = P_1 . T_2 / T_1$

P_2 = 210 x 150 /120

P_2 = 263 mmHg

Example 2: A cylinder contains a gas which has a pressure of 125 mmHg at a temperature of 200 K. Find the temperature of the gas which has a pressure of 100 mmHg if the volume remains constant.

Solution: P1 = 125 mmHg, T1 = 200 K, T2 = ?, P2 = 100 mmHg

$P_1 / T_1 = P_2 / T_2$

$T_2 = T_1 . P_2 / P_1$

T_2 = 200 x 100 / 125

T_2 = 160 K

The General Gas Equation

Now we can combine everything we have into one proportion:

The volume of a given amount of gas is proportional to the ratio of its Kelvin temperature and its pressure.
Same as before, a constant can be put in:

PV / T = C

As the pressure goes up, the temperature also goes up, and vice-versa.
Also same as before, initial and final volumes and temperatures under constant pressure can be calculated.

$P_1V_1 / T_1 = P_2V_2 / T_2$

Example 1: 500 liters of a gas are prepared at 1 atm and 200 °C. The gas is placed into a tank under high pressure. When the tank cools to 20.0 °C, the pressure of the gas is 30 atm. What is the volume of the gas?

Solution: V_1 = 500 L, P_1 = 1 atm, T_1 = 200 °C = (200 + 273), T_2 = 20 °C = (20 + 273), P_2 = 30 atm, V_2 = ?

$P_1V_1 / T_1 = P_2V_2 / T_2$

$V_2 = P_1V_1T_2 / P_2T_1$

$V_2 = 1 \times 500 \times 293 / 30 \times 473$

V_2 = 10.3 L

Example 2: What is the final volume of a 400 cm^3 gas sample that is subjected to a temperature change from 22 °C to 30 °C and a pressure change from 760 mm Hg to 360 mm Hg?

Solution: V_2 = ? V_1 = 400 cm^3, T_1 = 22 °C = (22 + 273), T_2 = 30 °C = (30 + 273), P_1 = 760 mmHg, P_2 = 360 mmHg

$P_1V_1 / T_1 = P_2V_2 / T_2$

$V_2 = P_1V_1T_2 / P_2T_1$

$V_2 = 760 \times 400 \times 303 / 360 \times 295$

V_2 = 867 cm^3

Example 3: What is the volume at STP of 720 mL of a gas collected at 20 °C and 3 atm pressure?

Solution: (At STP, T_2 = 273 °C, P_2 = 1 atm), V_2 = ?,V_1 = 720 mL, T_1 = 20 °C = (20 + 273), P_1 = 3 atm

$P_1V_1 / T_1 = P_2V_2 / T_2$

$V_2 = P_1V_1T_2 / P_2T_1$

$V_2 = 3 \times 720 \times 273 / 1 \times 293$

V_2 = 2013 mL

Dalton's Law of Partial Pressures

This law states that the total pressure P exerted by a mixture of gases say A, B, C and D is equal to the sum of the partial pressures of each constituent, provided the gases would not chemically react together. In other words, If more than one gas occupy a single container then the number of moles of each gas is in proportional to the pressure of each gas (the gas' partial pressure) and the total pressure is equal to the sum of all the partial pressures.

$P_{total} = P_A + P_B + P_C + P_D$

Where the small P's are the partial pressure of the individual gases. If a gas is collected over water, it likely to be saturated with water vapour and the total pressures become

$P_{total} = P_{gas} + P_{water\ vapour}$

$P_{gas} = P_{total} - P_{water\ vapour}$

The Ideal Gas Law

Ideal gas, or perfect gas, is the theoretical substance that helps establish the relationship of four gas variables, pressure (P), volume(V), the amount of gas(n)and temperature(T). All of these gas laws are based on "ideal" gases. Ideal gases have the following properties:

1. All gas molecules are in motion, and move randomly.
2. Each time the gas particles collide, kinetic energy is conserved (this is called elasticity).
3. The volume of the molecules of gas is negligible (meaning so small it's not worth stating).
4. Gases do not attract or repel other gas molecules.
5. The kinetic energy of a gas is directly proportional to its temperature (in Kelvin), and is the same for all gases at the same temperature.

Most gases found in nature do not meet all of these qualifications for ideal gas; however, they follow the rules closely enough that we can still use all of the equations above in theory, and it will not present a significant difference from what occurs in nature.

But over a wide range of temperature, pressure, and volume, real gases deviate slightly from ideal. Since, according to Avogadro, the same volumes of gas contain the same number of moles, chemists could now determine the formulas of gaseous elements and their formula masses. The idea gas law is:

$PV = nRT$

Where n is the number of moles of the number of moles and R is a constant called the universal gas constant and is equal to approximately 0.0821 L-atm / mole-K.

Example 1: At 655mm Hg and 25.0°C, a sample of Chlorine gas has volume of 750mL. How many moles of Chlorine gas at this condition?

Solution:

P = 655mm Hg (655/760 = 0.862 atm)

T = (25+273)K

V = 750mL=0.75L

n = ?

R = 0.0821 L-atm / mole-K

Solution

n=PV/RT

n = 0.862 x 0.75 / 0.0821 x 298

n = 0.026 mol

Example 2: A sample of butane (C_4H_{10}) of mass 3.728 g is placed in an evacuated bulb of volume 489 mL at 25 °C. What is its pressure?

Solution: M = 3.728 g, V = 489 mL = (0.489L), T = 25 °C = (25 + 273)K

PV = nRT

$$n = 3.728g \times \frac{1 \text{ mol}}{(4 \times 12.011 + 10 \times 1.008)\text{g/mol}}$$

$$n = 6.414 \times 10^{-2} \text{ mol}$$

P = nRT / V

= 6.414×10^{-2} x 0.0821 x 298 / 0.489

= 3.2 atm

Assessment

1. $P_1V_1 = P_2V_2$ supports
   a. Charles' law
   b. Boyle's law
   c. Graham's law
   d. Avogadro's law

2. Kelvin temperature can be converted into Celsius temperature by
   a. °C = K – 273
   b. K + 273
   c. °C + 273

K
d. K + 273
°C

3. From the ideal gas equation, PV = nRT, the unit of n is
   a. atm dm^3
   b. atm dm^3 K^{-1}
   c. moles
   d. K^{-1} $mole^{-1}$

4. How many atoms are contained in 2g of Hydrogen gas (H=1)
   a. 6.02×10^{23} atoms
   b. 6.02×10^{23} molecules
   c. 12.04×10^{23} atoms
   d. 12.04×10^{23} molecules

Answers

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

Week 9

Topic: Gas Laws

GAY LUSSAC'S LAW OF COMBINING VOLUMES

Gay Lussac's Law of Combining Volumes states that when gases react, they do so in volumes which bear a simple ratio to one another, and to the volume of the product(s) formed if gaseous, provided the temperature and pressure remain constant.

The law explains experimental facts about how gaseous atoms combine.

Example 1: For the reactions:

(i) $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

1 vol 3 vols 2 vols

1 volume of nitrogen combines with 3 volumes of hydrogen to form 2 volumes of ammonia.

(ii) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

2 vols 1 vol 2 vols

2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of steam.

(iii) $Cl_{2(g)} + H_{2(g)} \rightarrow 2HCl_{(g)}$

1 vol 1 vol 2 vols

1 volume of chlorine gas combines with 1 volume of hydrogen to form 2 volumes of hydrochloric acid.

Example 2: Consider the reaction: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

(a) What volume of steam is formed from 20 cm^3 of hydrogen and 20 cm^3 of oxygen mixed together? (b) What gas(s) is in excess, and by what amount?
Solution:

(a) The ratio of their volumes is 2 vols : 1 vol → 2 vols

20 vols : 10 vols → 20 vols

That means, 20 cm^3 of hydrogen will combine with 10 cm^3 of oxygen to form 20 cm^3 of steam.

(b) Oxygen is in excess by 10 cm^3.

AVOGADRO'S LAW: THE VOLUME AMOUNT LAW

Gives the relationship between volume and amount when pressure and temperature are held constant. According to Avogadro, equal volumes of different (ideal) gases at the same temperature and pressure contain equal numbers of molecules (moles) of the different gases. This law states that equal volume of all gases under the same temperature and pressure contain the same number of molecules.

For example, in the reaction

$2H_{2(g)}$ + $O_{2(g)}$ $2H_2O_{(g)}$

2 moles 1 mole 2 moles

2 vol 1 vol 2 volume

Volume Ratio of Reactants : Products :: 3Vol : 2Vol

If the amount of gas in a container is increased, the volume increases. If the amount of gas in a container is decreased, the volume decreases.

V n

As before, a constant can be put in:

$V / n = C$

This means that the volume-amount fraction will always be the same value if the pressure and temperature remain constant.

$V_1 / n_1 = V_2 / n_2$

Example 1: What volume of oxygen will be required to obtain 100 cm3 of sulphur (VI) oxide form sulphur (IV) oxide?

Solution:

	$So_2(g)$	+ $\frac{1}{2}O_2(g)$	———>	$SO_3(g)$
	1 mol	½ mol		1 mol
By Avogadro's law:	1 vol	½ vol		1 vol
	100 cm3	50 cm3		100 cm3

The volume of oxygen gas = 50 cm3

Example 2: A ballon with 4.00g of Helium gas has a volume of 500mL. When the temperature and pressure remain constant. What will be the new volume of Helium in the ballon if another 4.00g of Helium is added into the ballon?

Solution:

Using Avogadro's Law, $V_1 / n_1 = V_2 / n_2$

$V_1 = 500mL$

$V_2 = ?$

$n_1 = 4.00g / 4.00g/mol = 1$ mol

$n_2 = 8.00g / 4.00g/mol = 2$ mol

$V_2 = n_2 \cdot V_1 / n1$

$= 2 \times 500 / 1$

$V_2 = 1000$ mL or 1L

<u>GRAHAM'S LAW OF DIFFUSION</u>

Graham's law of diffusion states that the rate of diffusion of a gas, R is inversely proportional to the square root of its density d.

$$Rate_{diffusion} \propto \frac{1}{\sqrt{density}}$$

Since volumes of different gases contain the same number of particles, the number of moles per liter at a given T and P is constant. Therefore, the density of a gas is directly proportional to its molar mass (MM).

$$Rate_{diffusion} \propto \frac{1}{\sqrt{MM}}$$

Rate = $k / \sqrt{d}$ or $\sqrt{MM}$

Where K is the proportionality constant. If we compare two gases A and B with rate of diffusion R_A and R_B and d_A and d_B

$R_A / R_B = \sqrt{d_B} / \sqrt{d_A}$

Rate is the reciprocal of time t, i.e.

$R = {}^1/t$

By substituting, $R_A = {}^1/\ t_A$ and $R_B = {}^1/t_B$ we get

$R_A / R_B = t_B / t_A = \sqrt{d_B} / \sqrt{d_A}$

Density = ½ (Relative molecular mass of gas)

Generally:

$R_A / R_B = t_B / t_A = \sqrt{d_B} / \sqrt{d_A} = \sqrt{M_A} / \sqrt{M_B}$

Example 1: What is the relative rate of diffusion of NH_3 compared to He? Does NH_3 effuse faster or slower than He? If the helium in problem 3 takes 20.0 seconds to effuse, how long will NH_3 take?

Solution: (a)

$$\frac{r_{He}}{r_{NH_3}} = \sqrt{\frac{M_{NH_3}}{M_{He}}}$$

$$\frac{r_{He}}{1\ m/s} = \sqrt{\frac{17.031\ amu}{4.003\ amu}}$$

$$r_{He} = 2.06\, m/s$$

Helium will effuse faster at 2.06 times the rate of ammonia.

(b)

$$r_{NH_3} = 1m/s\ ,\ r_{He} = 2.06\, m/s$$

$$\frac{2.06\ m}{s}\ x\ 20.0\ s = 41.2\ m$$

$$41.2\ m\ x\ \frac{1\ s}{1\ m} = 41.2\ \ seconds$$

It will take ammonia 41.2 seconds to effuse the same distance.

Example 2: An unknown gas diffused 0.25 times faster than He. What is the molar mass of the unknown gas?

Solution:

$$\frac{r_{He}}{r_x} = \sqrt{\frac{M_x}{M_{He}}}$$

$$\frac{1\ m/s}{0.25\ m/s} = \sqrt{\frac{M_x}{4.003\ amu}}$$

$$\left(\frac{1\ m/s}{0.25\ m/s}\right)^2 = \left(\sqrt{\frac{M_x}{4.003\ amu}}\right)^2$$

$$16 = \frac{M_x}{4.003\ amu}$$

$$M_x = 64.05\ amu$$

Example 3: 200cm3 of hydrogen diffused through a porous pot in 40 seconds. How long will it take 300cm3 of chlorine to diffuse through the same pot?

Solution:

200cm3 of hydrogen diffused in 40 secs

300cm3 of hydrogen will diffuse in: 300 x 40 / 200

= 60s

Using the equation: $t_{H2} / t_{Cl2} = \sqrt{M_{H2}} / \sqrt{M_{Cl2}}$

Where t_{H2} = 60s, M_{H2} = 2, M_{Cl2} = 71, t_{Cl2} = ?

$t_{Cl2} = t_{H2} \times \sqrt{M_{Cl2}} / \sqrt{M_{H2}}$

t_{Cl2} = 60 x √71/2

= 60 x √35.5

= 60 x 5.96

= 357.5s

The time of diffusion of chlorine = 358s

Note: Diffusion – The rate at which two gases mix.

Effusion – The rate at which a gas escapes through a pinhole into a vacuum.

Assessment

1. If equal amounts of helium and argon are placed in a porous container and allowed to escape, which gas will escape faster and how much faster?
2. In demonstration of Graham's law fumes are produced at point of
   a. junction of gases
   b. where pressure is high
   c. where area is greater
   d. at end of tube
3. A 6.0 L sample at 25°C and 2.00 atm of pressure contains 0.5 mole of a gas. If an additional 0.25 mole of gas at the same pressure and temperature are added, what is the final total volume of the gas?
4. Number of atoms or molecule present in one molar mass of an entity is known as
   a. Bohr number
   b. Mass number
   c. Avogadro's number
   d. Atomic number
5. If the density of hydrogen is 0.090 g/L and its rate of diffusion is 5.93 times that of chlorine, what is the density of chlorine?

Answers

1. Set $rate_1$ = He = x
   Set $rate_2$ = Ar = 1
   The molecular weight of He = 4.00
   The molecular weight of Ar = 39.95
   Graham's Law is: $r_1 / r_2 = \sqrt{MM_2} / \sqrt{MM_1}$
   Substituting, we have:
   $x / 1 = \sqrt{(39.95 / 4.00)}$
   x = 3.16 times as fast.
2. A
3. First, express Avogadro's law by its formula: $V_i/n_i = V_f/n_f$
   where
   V_i = initial volume
   n_i = initial number of moles
   V_f = final volume
   n_f = final number of moles

For this example, V_i = 6.0 L and n_i = 0.5 mole. When 0.25 mole is added:
$n_f = n_i + 0.25$ mole
$n_f = 0.5$ mole = 0.25 mole
$n_f = 0.75$ moleThe only variable remaining is the final volume. $V_i/n_i = V_f/n_f$Solve for V_f
$V_f = V_in_f/n_i$
$V_f = (6.0 \text{ L} \times 0.75 \text{ mole})/0.5$ mole
$V_f = 4.5 \text{ L}/0.5$ $V_f = 9$ L

4. C

5. Set $rate_1 = H_2 = 5.93$
Set $rate_2 = Cl_2 = 1$
The molecular weight of H_2 = 2.02
The molecular weight of Cl_2 = x.
By Graham's Law:
$5.93 / 1 = \sqrt{(x / 2.02)}$
x = 71.03 g/mol
Determine gas density using the molar volume:
71.03 g / 22.414 L = 3.169 g/L

Week 10

Topic: Air

Air

The atmosphere surrounding the earth contains air. Air is a mixture of gases composed mainly of Nitrogen and Oxygen and small amounts of Carbon (IV) Oxide, noble gases and water vapour. *Air* is the general name for the mixture of gases that makes up the Earth's atmosphere. Air around us is a mixture of many gases and dust particles. It is the clear gas in which living things live and breathe. It has an indefinite shape and volume. It has no color or smell. It has mass and weight. It is a matter as it has mass and weight. Air creates atmospheric pressure. There is no air in the vacuum of the cosmos.

Nitrogen 78%, Oxygen 21%, Noble gases 1%, Carbon dioxide 0.03%, water vapour and dust.

Chemical Composition of Air

- Nearly all of the Earth's atmosphere is made up of only five gases: nitrogen, oxygen, water vapor, argon, and carbon dioxide. ...
- Nitrogen — N_2 — 78.084%
- Oxygen — O_2 — 20.9476%
- Argon — Ar — 0.934%
- Carbon Dioxide — CO_2 — 0.0314%
- Neon — Ne — 0.001818%
- Helium — He — 0.000524%
- Krypton — Kr — 0.000114%

Evidence that Air is a Mixture

1. The constituents of air can be separated by physical methods. For example, if highly compressed air is allowed to escape through a minute aperture, it expands suddenly and then loses heat to the environment. If this processes is continually repeated, air will be liquefied.
2. The constituents of air still retain their individual properties

3. If the appropriate proportions of each of the different constituents of air are mixed together under ordinary conditions, there will be no evidence of chemical combination such as heat production or volume change

4. The composition of air cannot be represented by a chemical formula

Properties of Air

1. Air has weight

2. Air can be compressed

3. Air takes up space

4. Air contains some water vapour

5. Air has velocity

Combustion of Substances in Air

Most substances burn in air. Burning or combustion is a chemical process which is frequently accompanied by the production of heat and light. Combustible material usually combine with atmospheric air for this reaction to occur. Example, when candle wax, a hydrocarbon is ignited, it melts, vapourizes and decomposes into its constituent elements namely hydrogen and carbon. These then combine with atmospheric oxygen during the process of burning to produce water and carbon (iv) oxide.

Flames

Flames are produced when substances burn. A flame can be described as a region where gases combine chemically with the production of heat and light. In most cases, it involves oxygen. Flames are not homogeneous but are composed of several defined zones. A flame maybe luminous or non-luminous. The luminosity of a flame is caused by the presence of solid particles in the flame. An increase in the temperature and pressure of the burning gases also increases luminosity.

Types of Flame

1. Hydrogen flame – Hydrogen burns with a very faint non-luminous flame. The structure of hydrogen flame is simple, consisting of only two regions; the unburnt gas zone and the zone of complete combustion

2. Candle flame – This burns with luminous flame. There are 4 zones
   – zone of unburnt gas around wick

– the bright yellow luminous zone where there is incomplete burning of hydrocarbon due to insufficient air supply
– the barely visible non luminous zone on the outside where complete combustion of carbon occurs due to sufficient air
– the blue zone at the base of the flame with a region of complete combustion

3. Bunsen flame – A Bunsen burner is built with an air inlet at the base of the burner so that a stream of air can be supplied to the flame together with the fuel gas. This supplements the external supply of air, and allows a more complete combustion of the fuel. The air-hole can on can also be adjusted. The fuel Bunsen burner is a mixture of hydrocarbon gases (methane and butane) and some hydrogen and carbon (ii) oxide. The products of combustion are water, carbon(iv) oxide and soot. To produce a luminous flame, the airhole should be closed. The flame produced is high and wavy with a large bright yellow zone but it is not hot. To produce a non-luminous flame, the air hole should be open. The flame is non-luminous but much hotter and cleaner. A Bunsen burner has 3 zones
– the unburnt gas zone
– the luminous zone
– the outer-most non-luminous zone which has increased in size because the flame has sufficient air

Assessment

1. Combustion is a chemical reaction which is always accompanied by
a. heat
b. heat and light
c. heat and energy
d. heat and power

2. All these are causes of luminosity in flames except
a. solid particles
b. solid particles and increased temperature
c. solid particles and increased pressure
d. size of materials burnt

3. These gases burn with blue flame except
a. H_2
b. CH_4
c. C_2H_2
d. CO

4. Which of these is not necessary for metallic corrosion
a. water
b. oxygen

c. heat
d. sulphur IV oxide

5. The most abundant noble gas in nature is
a. Neon
b. Radon
c. Helium
d. Argon

Answers

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

<u>S.S.S 1</u>

<u>CHEMISTRY</u>

<u>THIRD TERM</u>

TABLE OF CONTENT

Week 1

Topic: Acids

Acids have long been associated with sour taste of some fruits such as lime and lemon. Their ability to change litmus solution from blue to red is well known. There are two classes of acids – organic and inorganic acids. The former occurs as a natural product in plants and animal while the latter can be prepared from mineral elements or inorganic matter. An organic acid is an organic substance that has the properties of an acid, while inorganic acids, also known as mineral acids, come from inorganic substances. Some examples of inorganic acids include suphuric acid, hydrochloric acid, nitric acid, boric acid and hydrofluoric acid.

Mineral/Inorganic Acids

- They are generally much stronger
- Most do not occur naturally
- They usually have simpler molecules

Examples are:

- Sulphuric Acid
- Nitric Acid
- Hydrochloric Acid
- Phosphoric Acid
- Carbonic Acid

Organic Acids

- They naturally occur
- They are found in vegetables, fruit and other foodstuffs
- They are usually weaker and less corrosive

Examples are:

- Ethanoic Acid
- Citric Acid
- Lactic Acid

- Tartaric Acid
- Acetic Acid

ORGANIC ACID	**SOURCE**
Ethanoic acid	Vinegar
Lactic acid	Milk
Citric acid	Lime, lemon
Amino acids	Protein
Fatty acids	Fats and Oils
Ascorbic acid	Oranges
INORGANIC ACIDS	SOURCE
Hydrochloric acid HCl	Hydrogen and Chlorine
Tetraoxosulphate(vi) acid H_2SO_4	Hydrogen, Sulphur and Oxygen
Trioxonitrate(v) acid HNO_3	Hydrogen, Nitrogen and Oxygen

What is the difference between Organic and Inorganic Acids?

¤ Organic acids contain carbon, and inorganic acids don't contain carbon.

¤ Generally organic acids are weaker acids than inorganic acids.

¤ Most of the organic acids are insoluble in water (sometimes miscible with water), but soluble in organic solvents. But inorganic acids are generally well soluble in water and non soluble in organic solvents.

¤ Organic acids have a biological origin, whereas inorganic acids haven't. Inorganic acids are derived from inorganic compounds/mineral sources.

¤ Mineral acids are highly reactive with metals, and they have corrosive ability than the organic acids.

Assessment

Mention the sources of the following Organic acids

1. Lactic acid
2. Ascorbic acid
3. Fatty acids
4. Amino acid

Differentiate the following acids into Organic and Inorganic acids

1. Acetic acid
2. Citric acid
3. Sulphuric acid
4. Benzoic acid
5. Nitrous acid
6. Phosphoric acid
7. Taurine
8. Uric acid
9. Tartaric acid
10. Chromic acid
11. Phenol
12. Cinnamic acid
13. Fumaric acid

Week 2

Topic: Acids

Introduction

Acids form a class of chemical substances which contain hydrogen ions in aqueous solution, H^+ (aq), as the only positive ion. Acids are usually classified into mineral (Inorganic) or organic acids. An acid is a substance which produces hydrogen ions or protons as the only positive ion when dissolved in water. Acids dissolve in water to produce hydrogen ions H^+ as the only positive ions together with the corresponding negative ions. This process is known as ionization.

Acids in Solution

Acids are substances that form hydrogen ions when dissolved in water. A hydrogen ion is actually a proton. Therefore, acids are called the proton donors.

Strong Acids ionize completely in water to give hydrogen ions and anions. The concentration of hydrogen ions is very high in such acid solutions.

Examples:

$HCl \rightarrow H^+ + Cl^-$

$HNO_3 \rightarrow H^+ + NO_3^-$

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

The hydrogen ions produced will combine with the water molecule to form hydroxonium ions (H_3O^+)

$H^+ + H_2O \rightarrow H_3O^+$

Therefore the reaction can also be written as:

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

$H_2SO_4 + H_2O \rightarrow H_3O^+ + SO_4^{2-}$

The characteristic properties of an acid in solution are due to the presence of these hydrogen ions.

Weak acids are only partially ionized in water. Such acid solutions have a low concentration of hydrogen ions.

Examples:

$H_2CO_3 \leftrightarrow 2H+ + CO_3^{2-}$

$H_3PO_4 \leftrightarrow 3H^+ + PO_4^{3-}$

$CH_3COOH \leftrightarrow 2H^+ + SO_3^{2-}$

Basicity of Acids

Basicity of an acid is the number of replaceable hydrogen ions, H^+ in one molecule of an acid.

There are 3 common types of Basicity of an acid, these are:

- Monobasic (monoprotic)
- Dibasic (diprotic)
- Tribasic (triprotic)

Monobasic Acids

The monobasic acids are the acids that produced one H^+ ion from each acid molecule.

Examples:

$HCl \rightarrow H^+ + Cl^-$

$HNO_3 \rightarrow H^+ + NO_3^-$

$CH_3COOH \rightarrow CH_3COO^- + H^+$

Dibasic Acids

The dibasic acids are the acids that produced two H^+ ion from each acid molecule.

Examples:

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

$H_2SO_3 \rightarrow 2H^+ + SO_3^{2-}$

$H_2CO_3 \rightarrow 2H^+ + CO_3^{2-}$

Tribasic Acids

The tribasic acids are the acids that produced three H^+ ion from each acid molecule.

Example:

$H_3PO_4 \rightarrow 3H^+ + PO_4^{3-}$

Note: Acid containing more than three replaceable hydrogen ions in one molecule of the acid is called a polybasic acid

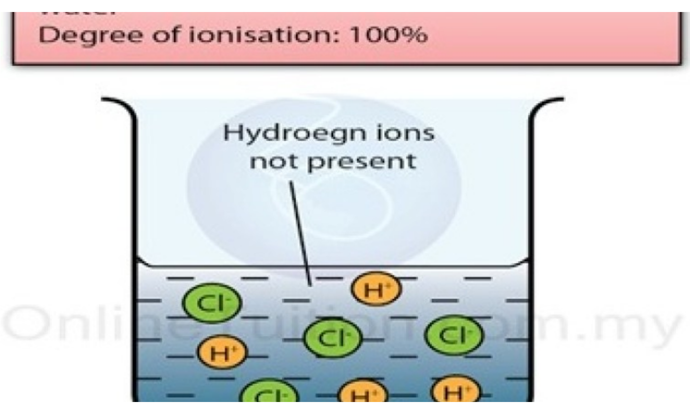
Strong Acids and Weak Acids

Acids are chemical substances that ionize/dissociate in the presence of water to produce hydrogen ions (or hydroxonium ions). Acids can be classified into 2 categories:

- Strong acids
- Weak acids

The strength of an acid depends on the degree of ionization/dissociation of the acid.

Strong Acids: Strong acids are acids that ionise completely to form hydrogen ions in water

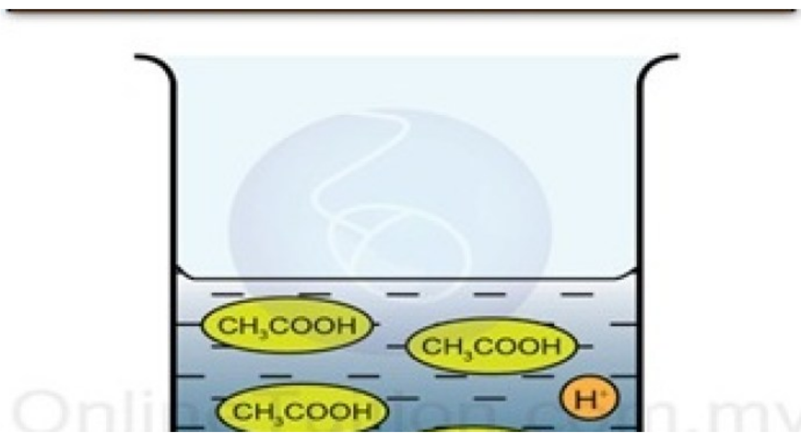


Examples:

- Sulphuric acid
- Hydrochloric acid

- Nitric acid

Weak Acids: Weak acids are acids that partly ionise to form hydrogen ions in water



Examples:

- Ethanoic acid
- Phosphoric acid
- Citric acid

Concentrated and Diluted Acids

- Concentrated Acid: This is one in which a large amount of the acid is dissolved in little amount of water
- Dilute Acid: This is one in which a little amount of the acid is dissolved in large quantity of water

Physical Properties of Acids

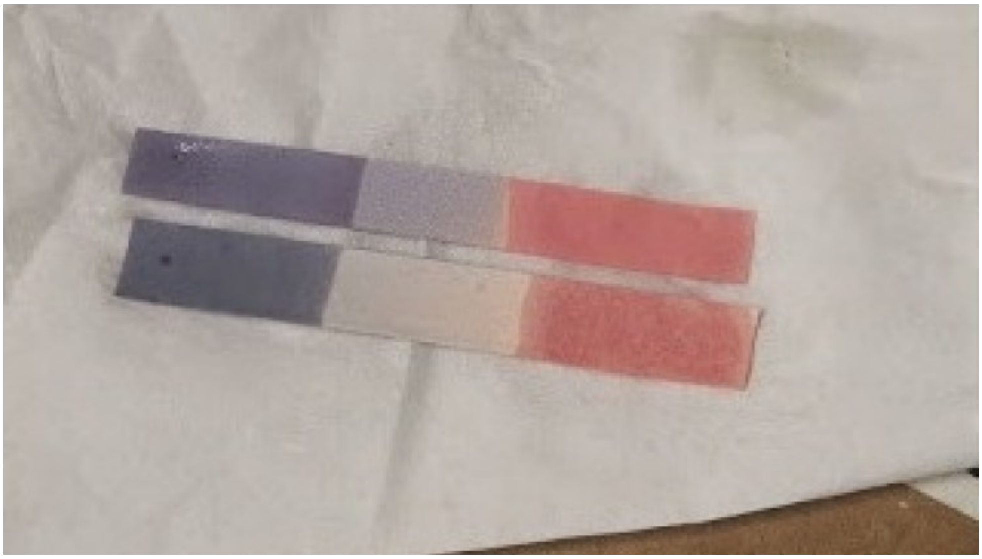
Acids have the following physical properties:

- Tastes sour
- Turns moist blue litmus to red
- pH value is less than 7 (i.e. < 7)
- Can conduct electricity
- Corrosive

Note:

Colour of Litmus in Acids

Litmus can be used as acid/alkali indicator. Image below shows the colour of litmus paper when immerse in acid and alkali. The litmus turn red in acids and turn blue in alkali.



pH value of acids

pH value is quantity to measure the concentration of hydrogen ions in a solution. The higher the concentration of hydrogen, the lower the pH value of the solution. Pure water has pH value of 7. All acids has pH value lower than 7.

Chemical Properties of Acids

Acids have the following chemical properties:

1. Acid + Reactive Metal → Salt + Hydrogen gas

Acids react with a metal that is more electropositive than hydrogen in the electrochemical series to produce salt and hydrogen gas. Acids do not react with copper and silver. This is a displacement reaction, where the metals that are placed above hydrogen in Electrochemical Series displace hydrogen from acid.

Acids + Reactive Metal → Salt + Hydrogen Gas

Examples:

Sulphuric acid + Zinc

$H_2SO_4 + Zn \rightarrow ZnSO_4 + H_2$

Nitric acid + Lead

$2HNO_3 + Pb \rightarrow Pb(NO_3)_2 + H_2$

2. Acid + Metal Carbonate → Salt + Water + Carbon Dioxide gas

Acids react with metal carbonates produces salt, water and carbon dioxide

Acids + Metal Carbonate → Salt + Water + Carbon Dioxide Gas

Examples:

Sulphuric acid + Lime Stone

$H_2SO_4 + CaCO_3 \rightarrow CaSO_4 + CO_2 + H_2O$

Nitric acid + Copper(II) Carbonate

$2HNO_3 + CuCO_3 \rightarrow Cu(NO_3)_2 + CO_2 + H_2O$

3. Acid + Base oxide → Salt + Water

Acids react with bases produces salt and water

Acids + Base Oxide → Salt + Water (Neutralisation)

Examples:

Sulphuric acid + Iron(II) Oxide

$H_2SO_4 + FeO \rightarrow FeSO_4 + H_2O$

Hydrochloric acid + Magnesium Oxide

$2HCl + MgO \rightarrow MgCl_2 + H_2O$

4. Acid + Alkali → Salt + Water

Acids react with alkali produces salt and water only. This is called a neutralisation reaction.

Acids + Alkali → Salt + Water (Neutralisation)

Examples:

Ethanoic Acid + Sodium Hydroxide

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

Hydrochloric Acid + Ammonia Solution

$HCl + NH_3 \rightarrow NH_4Cl$

Preparation of Acids

There are four methods which can be used to prepare acids. These are:

-
    - By dissolving acid anhydrides in water. Acid anhydrides are oxides of non-metals which will react with water to form the corresponding acids

Examples:

$CO_{2(g)} + H_2O_{(l)}$ ––> $H_2CO_{3(aq)}$

$SO_{2(g)} + H_2O_{(l)}$ ––> $H_2SO_{3(aq)}$

$SO_{3(g)} + H_2O_{(l)}$ ––> $H_2SO_{4(aq)}$

- By displacing a weaker more volatile acid form its sodium or potassium salt using a stronger but less volatile acid

Examples:

$NaCl_{(s)} + H_2SO_{4(aq)} \longrightarrow NaHSO_{4(aq)} + HCl_{(g)}$

$KNO_{3(aq)} + H_2SO_{4(aq)} \longrightarrow NaHSO_{4(aq)} + HNO_{3(g)}$

- By precipitating insoluble sulphide of metal from a solution of the metallic salt using hydrogen sulphide gas

Examples:

$(CH3COO)2Pb_{(aq)} + H_2S_{(g)} \longrightarrow PbS_{(s)} + 2CH_3COOH_{(aq)}$

$Cu(NO_3)_{2(aq)} + H_2S_{(g)} \longrightarrow CuS_{(s)} + 2HNO_{3(aq)}$

- By direct combination between the elements

$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$

$H_{2(g)} + Br_{2(g)} \longrightarrow 2HBr_{(g)}$

Uses of Acids

- Vinegar, used in the kitchen, is a liquid containing 3-6% acetic acid. It is used in pickles and in many food preparations.
- Lemon and orange juice contains citric acid. Citric acid is used in the preparation of effervescent salts and as a food preservative.
- Acids have been put to many uses in industry. Nitric acid and sulphuric acid are used in the manufacture of fertilizers, dyes, paints, drugs and explosives.
- Sulphuric acid is used in batteries, which are used in cars, etc. Tannic acid is used in the manufacture of ink and leather.
- Hydrochloric acid is used to make aqua regia, which is used to dissolve noble metals such as gold and platinum. It is also used to remove rust and used by industries to make chemicals.
- Sulphuric acid is used in manufacturing fertilizers such as super phosphate, ammonium sulpahte etc.
- Boric acid is used as a germicide or mild antiseptic
- Fatty acids is used in the manufacture of soaps via saponification
- Tartaric acid is used in making baking soda, soft drinks and health salts

Assessment

1. Give examples of the following acids – Monobasic, Dibasic and Tribasic acids
2. Mention 3 uses of Acids
3. Mention 3 physical properties of Acids

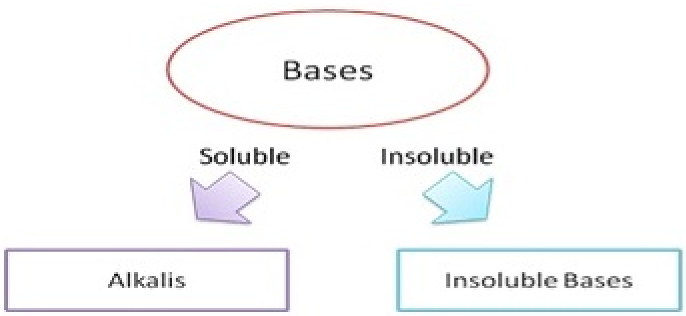
Week 3

Topic: Bases and Alkalis

Introduction

Bases are substances that, in aqueous solution, are slippery to the touch, taste bitter, change the colour of indicators (e.g., turn red litmus paper blue), react with acids to form salts, and promote certain chemical reactions (*base* catalysis). Bases are usually metallic oxides or metallic hydroxides.

Bases that are soluble in water are called alkalis.



In aqueous solution, alkali produces hydroxide ions (OH^-). In short, alkalis are substances that form hydroxide ions ($OH^-(aq)$) in water

Examples

Sodium hydroxide NaOH gives $Na^+(aq)$ and $OH^-(aq)$ ions,

$NaOH \rightarrow Na^+ + OH^-$

Calcium hydroxide $Ca(OH)_2$ gives $Ca^{2+}(aq)$ and $2OH^-(aq)$ ions.

$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$

Ammonia give NH_4^+ and OH^-

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

[Note: An alkali is a base soluble in water.]

In alkaline solution there are more OH^- ions than H^+ ions

Strength of Alkalis

Similar to strength of acids, the strength of an alkali is defined by its ability to ionise and release hydroxide ions (OH^-) in the solution.

In a solution of strong alkali, all the alkali molecules are ionised in the water to produce hydroxide ions

In a solution of weak alkali, only small portion of the molecules are ionised to release hydroxide ions.

Table below shows some example of strong/weak alkalis.

Alkali	
Strong	Weak
NaOH KOH LiOH	NH_3

Physical Properties of Alkali

The following are the physical properties of alkali

1. Alkalis are bitter in taste.
2. Alkalis turn litmus from red to blue.
   Like acid, alkali can change the colour of litmus. In alkali solution, the colour of litmus turn blue.
3. Alkalis are soapy to touch.
4. Alkalis has pH value more than 7
   pH value is a measurement of the concentration of hydrogen ions in a solution. Alkali has very low concentration of solution hydrogen ion, even lower than water. Hence the pH value of alkali is higher than 7. (Note: The pH value of water is 7. The lower the concentration of hydrogen ions, the higher the pH value.
5. Alkalis can conduct electricity
   When a base dissolve in water, it will dissociate and form hydroxide ions. The present of the freely move ions make alkali an electrolyte.

Chemical Properties of Bases/Alkalis

Alkalis react with acids to form a salt and water — this is a neutralisation reaction:

Reaction between Alkalis and Acids

Acid + Alkali → Salt + Water

Example:

Potassium hydroxide + Nitric Acid → Potassium Nitrate + Water

$KOH + HNO_3 \rightarrow KNO_3 + H_2O$

Alkali heat with Ammonium Salts

Alkalis, when warmed with ammonium salts, give off ammonia gas

Ammonium Salt + Alkali → Salt + Ammonia + Water

Example:
Ammonium Chloride + Sodium Hydroxide → Sodium chloride + Water + Ammonia

$NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$

More examples:

$2NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O + 2NH_3$

$(NH_4)_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O + 2NH_3$

<u>Preparation of Bases and Alkalis</u>

There are different ways of preparing bases and alkalis

- By burning metals in air or oxygen

$2Mg (s) + O_2(g) \longrightarrow 2MgO(s)$

$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$

- Sodium hydroxide is prepared by

1. Heating slaked lime with dilute sodium trioxocarbonate (IV) solution

$Ca(OH)_2(s) + Na_2CO_{3(aq)} \longrightarrow CaCO_{3(s)} + 2NaOH_{(aq)}$

2. By electrolysis of brine

- Ammonium hydroxide (NH_4OH) is prepared by heating a mixture of ammonium salt and slaked lime to produce ammonia gas which is then dissolved in water

$Ca(OH)_{2(s)} + (NH_4)_2SO_{4(s)} \longrightarrow CaSO_{4(s)} + 2H_2O_{(g)} + 2NH_{3(g)}$

$NH_{3(g)} + H_2O_{(l)} \longrightarrow NH_4OH_{(aq)}$

- Insoluble bases are prepared by method of precipitation. This is done by addition of sodium hydroxide or potassium hydroxide to soluble salt solution

$CuSO_{4(aq)} + 2KOH_{(aq)} \longrightarrow K_2SO_{4(aq)} + Cu(OH)_{2(s)}$

Uses of Bases

- Sodium hydroxide (caustic soda) is used in the manufacture of soap. It is used in petroleum-refining; in making medicines, paper, pulp, etc. It is used in making rayon.
- Calcium hydroxide is also known as slaked lime. It is used to neutralize acid in water supplies; in the manufacture of bleaching powder; as a dressing material for acid burns; as an antidote for food poisoning; in the preparation of fungicides and in the mixture of whitewash. It is mixed with sand and water to make mortar which is used in the construction of buildings. It is also used by farmers on the fields to neutralize the harmful effects of acid rain.
- Ammonium hydroxide is used to remove ink spots from clothes and to remove grease from window-panes. It is used in the cosmetic industry.
- Alkalis are used in alkaline batteries. Generally, potassium hydroxide is used in such batteries.
- Magnesium hydroxide is used in toothpaste to neutralise acid on teeth. It is also used in antacids to relieve indigestion

Assessment

1. A …….. is a substance which will neutralize an Acid to give salt and water only
2. …. turns red litmus paper blue
3. ………. is the process in which an acid reacts completely with an appropriate amount of an alkali to produce water and salt

4. The equation below represents
   $H^{+}_{aq} + OH^{-}_{aq} \rightarrow H_2O_l$
   a. Hydrolysis
   b. Hydration
   c. Neutralization
   d. Electron affinity

Answers

1. Base
2. Alkali
3. Neutralization
4. C

Week 4

Topic: Salts

Introduction

A salt is an ionic compound formed when the hydrogen of an acid is partly or completely replaced by a metal ion or ammonium ion. All salts are chemically and electrically neutral.
Example:

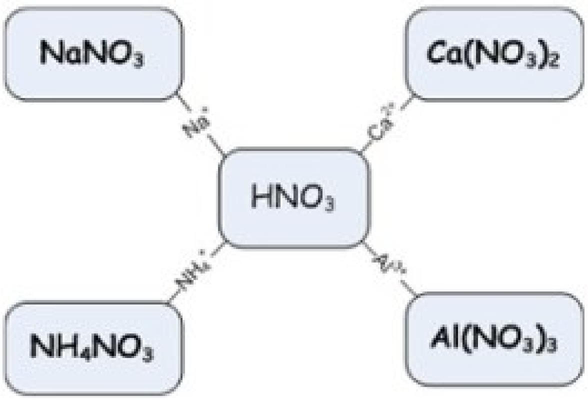


Diagram above shows that when the hydrogen ion in nitric acid is replaced by Na^+, Ca^{2+}, NH_4^+ or Al^{3+} ions, salts are formed.

Other examples:
Barium nitrate, zinc sulphate and tin nitrate are salts

There are 4 types of salt, these are:

- Nitrate
- Chloride
- Sulphate
- Carbonate

Classification of Salt

Salts are classified into four different types:

- Normal salts
- Double salts
- Mixed salts

- Complex salts

Normal Salts: Salts which produce one simple cation and one simple anion in aqueous solution are called normal salts.

The ions present in simple salt can be tested easily. Based on the nature of ions produced they are further classified into:

1. Neutral salts
2. Acidic salts
3. Basic salts

Neutral salts: Salt which is formed by complete neutralization of strong acid and base or weak acid and weak base is called neutral salt and it neither produces H^+ or OH^- in solution

Examples:

- NaCl (formed by neutralization of NaOH and HCl)

$HCl_{(aq)} + NaOH_{(aq)}$ —–> $NaCl_{(aq)} + H2O_{(l)}$

- K_2SO_4(formed from KOH and H_2SO_4)

$H_2SO_{4(aq)} + 2KOH_{(aq)}$ —–> $K_2SO_{4(aq)} + 2H_2O_{(l)}$

Acidic salts: Salt formed by partial neutralization of poly basic acid with a base. Acidic salt produces H^+ in solution.

Example:

- $NaHSO_3$(formed When poly basic acid H_2SO_3 is partially neutralized by NaOH)

$NaOH_{(aq)} + H_2SO_{3(aq)}$ ——–> $NaHSO_{3(aq)} + H_2O_{(aq)}$ (Acidic salt formed by partial neutralization)

$2NaOH_{(aq)} + H_2SO_{3(aq)}$ ———> $Na_2SO_{3(aq)} + 2H_2O_{(l)}$ (Neutral salt formed by complete neutralization)

Basic salt: Salt which is formed by partial neutralization of poly acidic base ($Ca(OH)_2$,$Fe(OH)_3$ etc.) with an acid. Such a salt produces OH^- ion in solution

Example:

- Ca(OH)Cl (formed by partial neutralization of $Ca(OH)_2$ with HCl)

$Ca(OH)_{2(aq)} + HCl_{(aq)}$ ——–> $Ca(OH)Cl_{(aq)} + H_2O_{(l)}$ (Basic salt)

$Ca(OH)_{2(aq)}$ + $2HCl_{(aq)}$ ——> $CaCl_{2(aq)}$ + $H_2O_{(l)}$ (Neutral salt)

Double salt: A salt formed from two different salts and whose solution gives test for all the ions present in it. In other words, double salt contains two different positive metallic ions and common negative acid radical or a positive metallic ion and ammonium ion and a common negative acid radical

Examples:

- $FeSO_4 (NH_4)_2SO_4 .6H_2O$ – Ammonium Iron (III) tetraoxosulphate (VI) or Mohr salt
- $K_2SO_4 Al_2(SO_4)_3 .12H_2O$ – Potassium Aluminium tetraoxosulphate (VI) dodecahydrate (Potash alum)

Mixed salt: When an acid is simultaneously neutralized by two bases or when a base is neutralized by two acids. They produce two cations or two anions and one cation.

Example: Ca(OCl)Cl -bleaching powder

Complex salt: Salt which produces one simple ion and a complex ion in aqueous solution. A complex salt does not answer the ions present in complex ion.

Example: $K_4(Fe(CN)_6$ – Potassium hexacyanoferrate (II)

Uses of Salts

Salts are used in the manufacture of many industrial, agricultural and consumer substances like chlorine gas, fertilizers and laxatives. They are also used as food preservatives, drying agents and anti-freeze.

Salt Hydrolysis

Hydrolysis is the reaction of salt with water to form a solution which is either acidic or basic.

Solubility of Salts

Solubility is the ability of a compound to dissolve in a solvent.

Table below shows the solubility of the salts of nitrate, sulphate, chloride and carbonate.

Salt	Solubility
Salt of potassium, sodium and ammonium	All soluble in water
Salt of nitrate	All soluble in water
Salt of sulphate	Mostly soluble in water except: (Pb) Lead sulphate

	(Ba) Barium sulphate (Ca) Calcium sulphate
Salt of chloride	Mostly soluble in water except: (Pb) Lead chloride (Ag) silver chloride (Hg) mercury chloride
Salt of carbonate	Mostly insoluble in water except: Potassium carbonate Sodium carbonate Ammonium carbonate

Notes: Lead halide such as lead(II) chloride ($PbCl_2$), lead(II) bromide ($PbBr_2$) and lead (II) iodide (PbI_2) are insoluble in cold water but soluble in hot water.

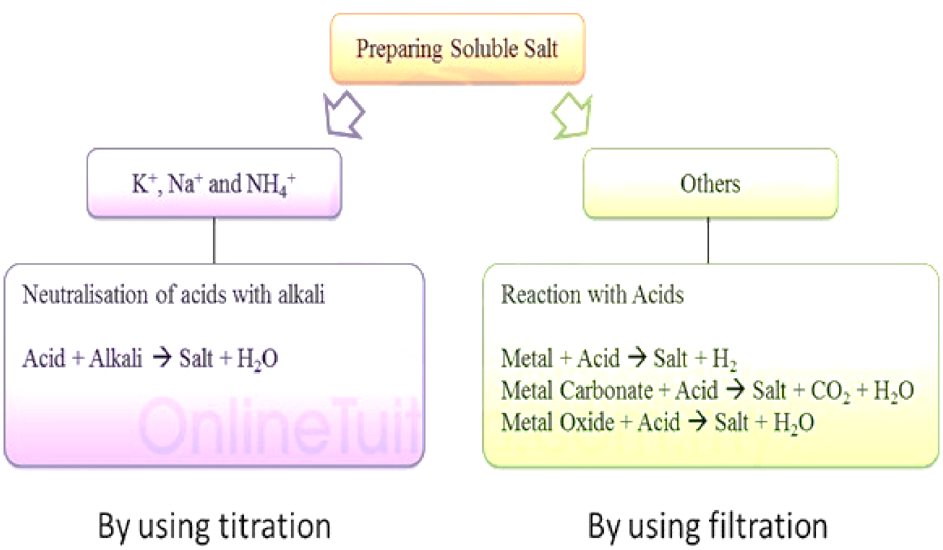
Preparation of Soluble Salts

There are 2 things to be considered when preparing a salt:

- What are the chemical used?
- How to separate the salt from other substance?

Method used to prepare salt depends on the solubility of the salt. Soluble salts are prepared from the reactions between an acid with a metal/ base/ metal carbonate.

Diagram below shows the chemical reaction that can be used to prepare the soluble salts.



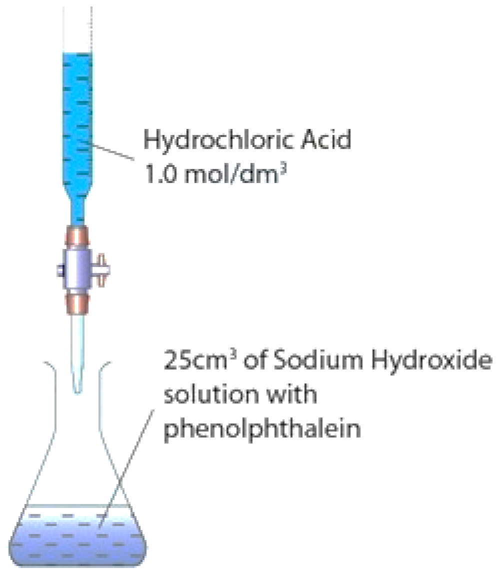
Preparing Salts of Potassium, Sodium and Ammonium

Potassium, sodium and ammonium salts are usually prepared through the reactions of acids with alkalis. Reacting acid with alkali will produce salt and water. The salt is prepared by titration method of acid and alkali using an indicator.

Acid + Alkali → Salt + Water

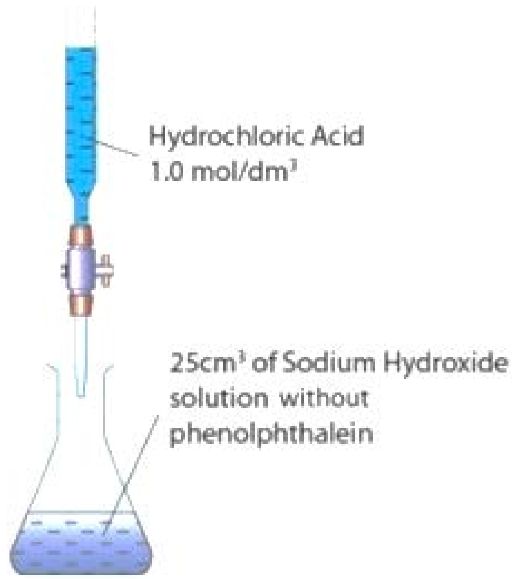
Steps to Prepare the Salts of Potassium, Sodium and Ammonium through Titration

Step 1 – Titration to Find the End Point



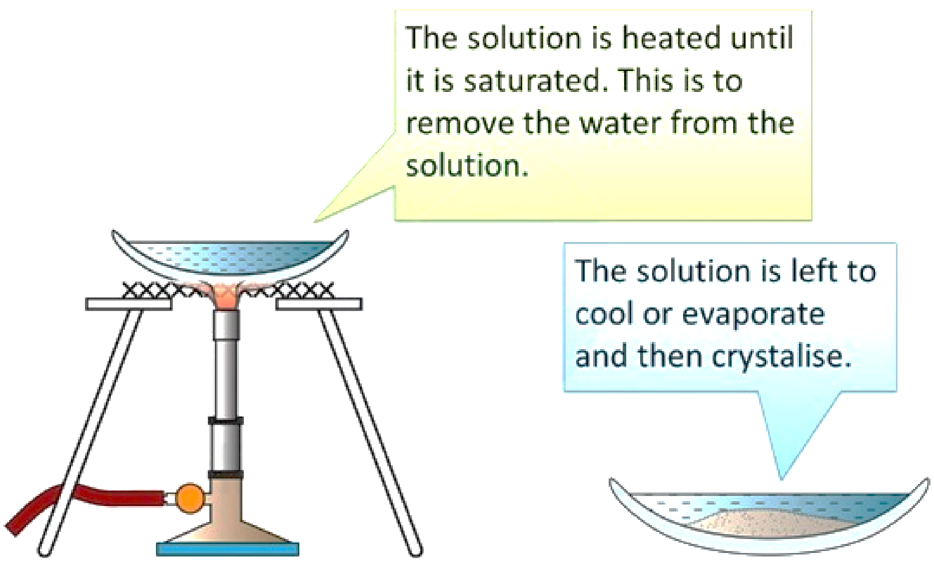
The end point is the point in a titration at which the two reactants have completely reacted. It is often marked by a colour change.

Step 2 – Titrate Without Indicator

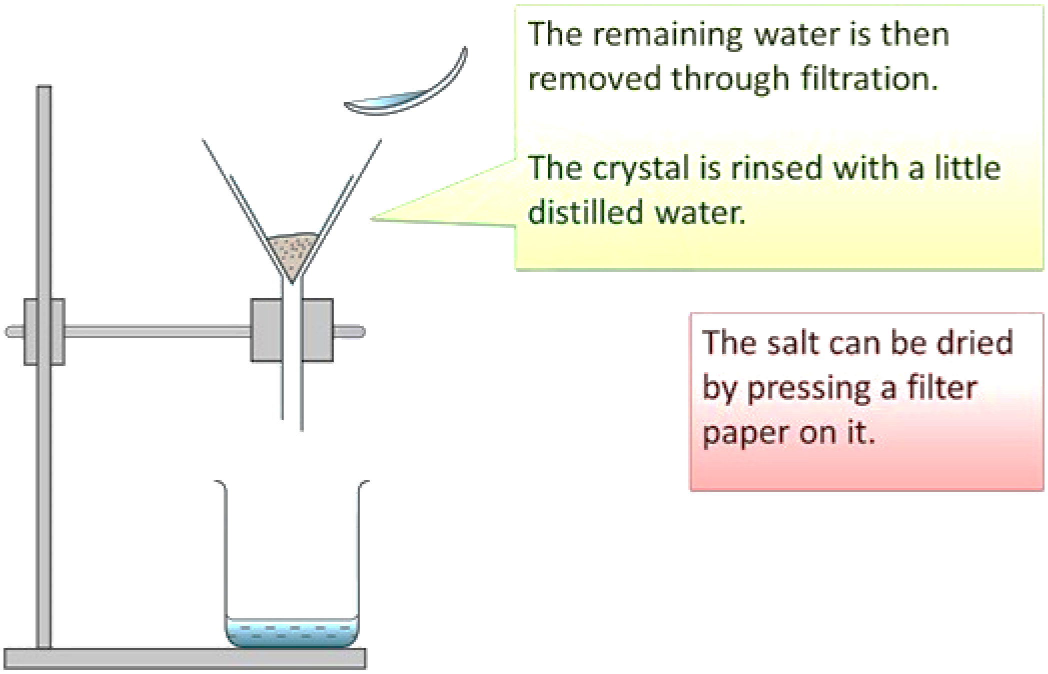


The product obtain in step 1 is contaminate by the indicator. The reaction is repeated by using the same amount of reactants as in step 1, without using any indicator.

Step 3 – Crystalisation



Step 4 – Filtration and Drying



Preparing Salts of Non-“Potassium, Sodium and Ammonium”

The salt non-potassium, sodium and ammonium is prepared by reacting acid with insoluble metal/metal oxide/metal carbonate:

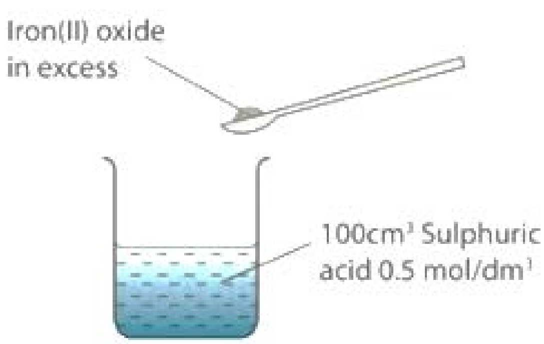
Acid + Metal Salt + Hydrogen (Displacement reaction)

Acid + Metal oxide Salt + Water (Neutralisation Reaction)

Acid + Metal carbonate Salt + Water + Carbon Dioxide

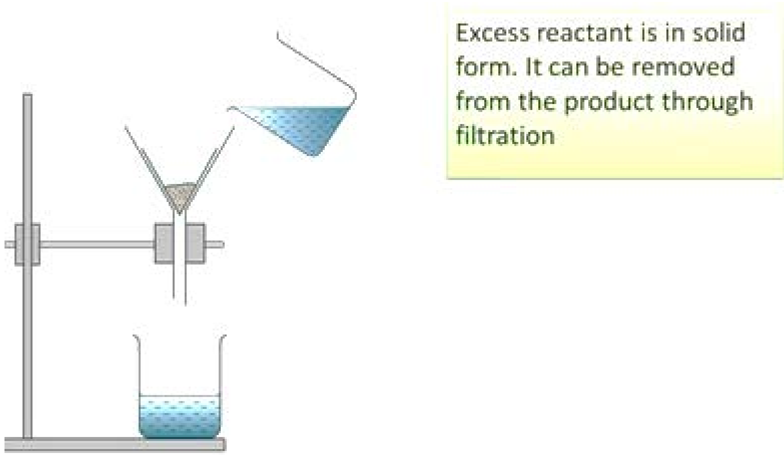
Below is the steps in preparing the soluble non-potassium, sodium and ammonium salts

Step 1 – The Reaction



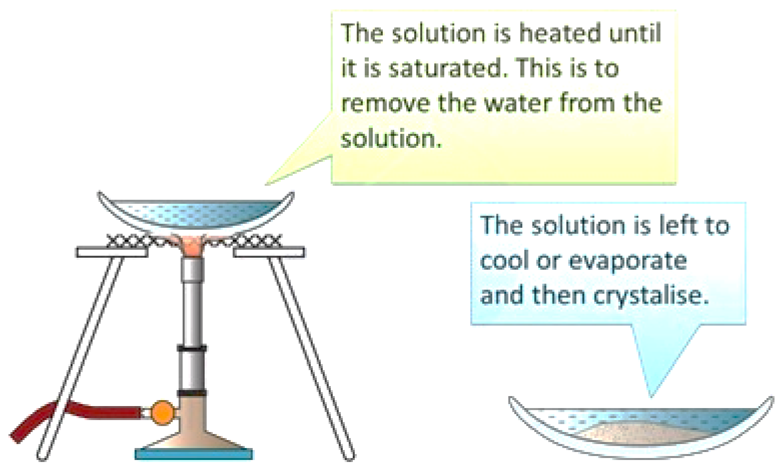
Add metal/metal oxide/metal carbonate powder until excess into a fixed volume of the heated acid

Step 2 – Filtration 1 to Remove Excess Reactant



Filter the mixture to remove excess metal/metal oxide/metal carbonate

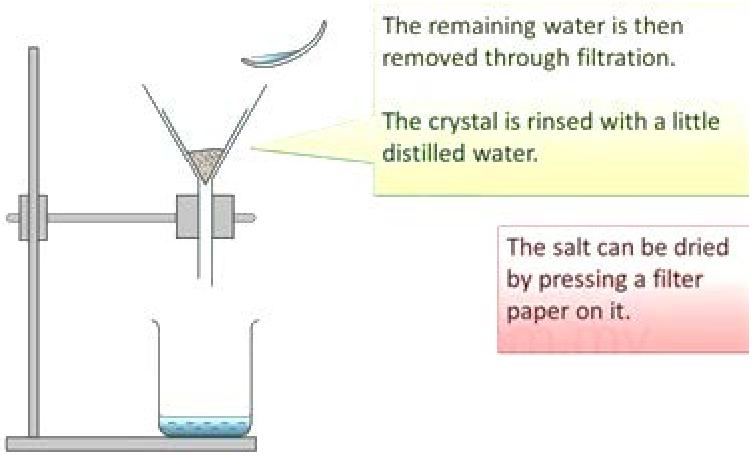
Step 3 – Crystallisation



Evaporate the filtrate until it becomes a saturated solution

Dip in a glass rod, if crystals are formed, the solution is saturated.

Step 4 – Filtration 2 to Collect the Solid Salt



Cooled at room temperature

Filter and dry the salt crystals by pressing them between filter papers

Chemical equation(s) for the reaction that can be used to prepare the following salts:

- Sodium Chloride

$NaOH + HCl \rightarrow NaCl + H_2O$

- Ammonium Nitrate

$NH_3 + HNO_3 \rightarrow NH4NO_3$

- Potassium sulphate

$KOH + H_2SO_4 \rightarrow K_2SO_4 + H_2O$

- Zinc Sulphate

$ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

$ZnCO_3 + H_2SO_4 \rightarrow ZnSO_4 + H_2O + CO_2$

- Lead(II) nitrate

$PbO + 2HNO_3 \rightarrow Pb(NO_3)_2 + H_2O$

$Pb + 2HNO_3 \rightarrow Pb(NO_3)_2 + H_2$

$PbCO_3 + 2HNO_3 \rightarrow Pb(NO_3)_2 + H_2O + CO_2$

- Copper sulphate

$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

$CuCO_3 + H_2SO_4 \rightarrow CuSO_4 + H_2O + CO_2$

Preparing Insoluble Salts

Insoluble salts can be made by ionic precipitation (is also called double decomposition/double displacement). This involves mixing a solution that contains its positive ions with another solution that contains its negative ions.
The equation of the reaction that can be used to prepare the following salt:

- Calcium sulphate

$CaCl_2 + NaSO_4 \rightarrow CaSO_4 + 2NaCl$

$Ca(NO_3)_2 + ZnSO_4 \rightarrow CaSO_4 + Zn(NO_3)_2$

- Lead chloride

$Pb(NO_3)_2 + 2NaCl \rightarrow PbCl_2 + 2NaNO_3$

- Copper carbonate

$CuSO_4 + Na_2CO_3 \rightarrow CuCO_3 + Na_2SO_4$

$CuCl_2 + K_2CO_3 \rightarrow CuCO_3 + 2KCl$

$Cu(NO_3)_2 + Na_2CO_3 \rightarrow CuCO_3 + 2NaNO_3$

Efflorescent, Deliquescent and Hygroscopic Substances

These terms are used to describe the behaviour of substances when exposed to the atmosphere. These substances either give up water to the atmosphere or absorb water from the atmosphere

Efflorescent compounds: These are hydrated salts that lose part or all their water of crystallisation to form a lower hydrate salt when exposed to atmosphere. Examples are washing soda i.e. sodium trioxocarbonate (IV) decahydrate ($Na_2CO_3.10H_2O$) which changes to monohydrate $Na_2CO_3.H_2O$

Equation: $Na_2CO_3.10H_2O \longrightarrow Na_2CO_3.H_2O + 9H_2O$

Another example is sodium tetraoxosulphate (VI) decahydrate known as Glawber's salt. It changes to the anhydrous salt on exposure to air for some days.

$Na_2SO_4.10H_2O \longrightarrow Na_2SO_{4(aq)} + 10H_2O$

Deliquescent compounds: These are compounds which when exposed to the atmosphere will absorb moisture and dissolve in it to form solutions

Examples are:

- Sodium hydroxide pellets
- Iron (III) chloride
- Phosphorous (V) oxide
- Potassium hydroxide
- Calcium chloride

Hygroscopic Compounds: These are compounds which when exposed to the atmosphere will absorb moisture, but will not dissolve in it and will only become sticky

Examples are:

- Copper (II) oxide
- Quicklime (Calcium oxide)
- Sodium trioxonitrate (V)

The only liquid hygroscopic compound is concentrated tetraoxosulphate (VI) acid. It will absorb moisture from the atmosphere and then becomes dilute

Assessment

1. .... are hydrated salts that lose part or all their water of crystallisation to form a lower hydrate salt when exposed to atmosphere.
2. .... are compounds which when exposed to the atmosphere will absorb moisture, but will not dissolve in it and will only become sticky
3. A .... is a compound formed when all or part of the ionizable hydrogen of an acid is replaced by metallic or ammonium salts
4. .... salts are formed by double decomposition and direct combination
5. These are types of salts except
   a. normal, acid, basic and complex salt
   b. normal, double, acid salt
   c. complex, basic, acid and hydrolysed salts
   d. double, basic and complex salt
6. Water in Crystalline salts provide
   a. hydrated bonds and colour
   b. colour and solubility base
   c. colour and shape
   d. shape and crystal lattice

7. Which of these is not a hygroscopic salt
   a. Calcium oxide
   b. Magnesium chloride
   c. Copper II oxide
   d. Sodium trioxocarbonate

Answers

1. Efflorescent Compounds
2. Hygroscopic Compounds
3. Salt
4. Insoluble
5. C
6. C
7. B

Week 5

Topic: Alkalinity and Acidity

Content

1. Alkalinity and Acidity
2. Calculations of pH and pOH
3. Indicators
4. Hygroscopy
5. Drying Agents

Alkalinity and Acidity

Alkalinity and Acidity are measured using a scale of numbers from 0 to 10 called the pH scale. A solution with pH value of 7 is neutral. A solution with a pH value less than 7 is acidic while a solution with pH value greater than 7 is alkaline. Acidity increases with decreasing pH while alkalinity increases with increasing pH.

Dissociation of water

Water is neutral and it ionizes very slightly to yield an equal number of hydrogen ions and hydroxide ions.

$[H^+] = [OH^-] = 1 \times 10^{-7}$ mol dm^{-3} (at 25°C)

The product of the two ionic concentrations gives the ionic product of water. It is represented by

$K_w = [H^+] = [OH^-] = 1 \times 10^{-7}$ (mol dm^{-3})^2 (at 25°C)
$= 10^{-14}$ mol^2 dm^{-6} (at 25°C)

K_w is constant under al circumstances at 25°C.

PH Scale

The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration to the base 10. It can also be defined as the measure of the concentration of hydrogen ions in solution or the measure of the alkalinity or acidity of a solution. Example, if the hydrogen ion concentration of a given aqueous medium is 10^{-5} mol dm^{-3} the acidity of the solution could be written in terms of its pH as follows

$[H^+] = 10^{-5}$

$\text{Log } [H^+] = \text{Log } 10^{-5}$

$= -5$

$pH = -\text{Log } [H^+] = -(-5) = 5.$

$pOH = 14 - pH = 14 - 5 = 9$

Since the concentration of OH ions is always inversely proportional to the H ions concentration, the pH value would indicate both the acidity and the alkalinity of the solution.

$[H^+] [OH^-] = 10^{-14}$

pH + pOH = 14 where pOH is the hydroxide ion index

A high pH value would indicate low hydrogen ion concentration i.e. weak acidity and a correspondingly high hydroxide ion concentration i.e. a strong alkalinity.

Since

$[H^+] = [OH^-] = 10^{-7}$

At neutrality, the neutral pH value is 7. A pH value which is less than 7 indicates an acidic medium and pH higher than 7 indicates an alkaline medium.

Example

Find the hydrogen and hydroxide ion concentrations in 1. 0.01 M tetraoxosulphate (vi) acid 2. 0.001 M potassium hydroxide solution.

a. tetraoxosulphate (vi) acid ionizes completely in solution. It is dibasic, so one mole would ionize to give 2 moles of hydrogen ions.

$H_2SO_{4\,(aq)} \rightarrow 2H^+_{\,(aq)} + SO_4^{2-}{}_{(aq)}$

In a 0.01 M solution,

$[H^+] - 2 \times 0.01 = 2 \times 10^{-2} \text{ mol dm}^{-3}$

$2 \times 10^{-2} \times [OH^-] = 10^{-14}$

$[OH^-] = 0.5 \times 10^{-12} \text{ mol dm}^{-3}$

$= 5 \times 10^{-13} \text{ mol dm}^{-3}$

b. $KOH_{\,(aq)} \rightarrow K^+ OH^-$

$[OH^-] = 10^{-3} \text{ mol dm}^{-3}$

$[H^+] [OH^-] = 1 \times 10^{-14}$ mol dm^{-3}

$[H^+] = \underline{1 \times 10^{-14}}$
1 x 10-3

$= 1 \times 10^{-11}$ mol dm^{-3}

Find the $[OH^-]$ given the pH or pOH. You are given that the pH = 4.5.

pOH + pH =14
pOH + 4.5 = 14
pOH = 14 – 4.5
pOH = 9.5

$[OH^-] = 10^{-pOH}$
$[OH^-] = 10^{-9.5}$
$[OH^-] = 3.2 \times 10^{-10}$ M

Find the hydroxide ion concentration of a solution with a pOH of 5.90.

pOH = -log$[OH^-]$
5.90 = -log$[OH^-]$
Because you're working with log, you can rewrite the equation to solve for the hydroxide ion concentration:

$[OH^-] = 10^{-5.90}$
To solve this, use a scientific calculator and enter 5.90 and use the +/- button to make it negative and then press the 10^x key. On some calculators, you can simply take the inverse log of -5.90.

$[OH^-] = 1.25 \times 10^{-6}$ M

Find the pOH of a chemical solution if the hydroxide ion concentration is 4.22×10^{-5} M.

pOH = -log$[OH^-]$
pOH = -log$[4.22 \times 10^{-5}]$

To find this on a scientific calculator, enter 4.22 x 5 (make it negative using the +/- key), press the 10^x key, and press equal to get the number in scientific notation. Now press log. Remember your answer is the negative value (-) of this number.
pOH = – (-4.37)
pOH = 4.37

Relationship Between pH and pOH

The pH and pOH of a water solution at 25°C are related by the equation below.

pH + pOH = 14

If either the pH or the pOH of a solution is known, the other can be quickly calculated.

Importance of pH

1. In the body, the acidic medium is required for digestion of food in the small intestine. For our bodies to function normally, the blood pH should be 7.4 .
2. Plants grow well in soils with pH of 7 or 8. Most soil pH vary from 4 – 9.
3. pH values are important in medicine, water treatment, sewage treatment etc

Indicators

Indicators are weak organic acids or bases which will produce different colours in solution according to the hydrogen ion concentration in that solution. An indicator is any substance that gives a visible sign, usually by a colour change, of the presence or absence of a threshold concentration of a chemical species, such as an **acid** or an alkali in a solution. An example is the substance called **methyl** yellow, which imparts a yellow colour to an alkaline solution. They ionize slightly in water. Thus the colour of the indicator is dependent on the relative proportion of ions and the molecules. This is determined by the degree of dissociation of the indicator, which in turn is dependent on the hydrogen ion concentration or pH of the medium. Generally, the complete change from one colour to another occurs gradually over two pH units. An indicator such as phenolphthalein will ionize to produce hydrogen ions together with negatively charged indicator ions.

Acid – Base Indicators

They are dyes which change colour according to the pH of the medium. Litmus is a common indicator which is red in acid and blue in alkali. It changes from purple to blue over a pH range of 5.0 to 8.0. Each indicator has its own specific pH range over which it changes colour.

Measuring pH of a Solution

The pH of a solution can be measured by Universal indicators and pH meters.

A universal indicator is made up of a mixture of various indicators which work at different ranges. By a series of successive colour changes, it can indicate pH values from about 3 to 11. These changes can be easily determined by comparing the colour obtained with that of the standards given.

A pH meter is used to measure hydrogen ion concentration. The instrument consists of 2 electrodes, a glass electrode and a reference electrode. A meter measure the voltage difference between the two electrodes and gives a pH reading. The pH of a solution can be measured accurately by using a pH meter. It can also measure the pH of dilute solutions.

Reaction of Substances When Expose to Air

When certain compounds are exposed to air, they either lose their water of crystallization or they absorb moisture from their surroundings. The terms efflorescence, deliquescent and hygroscopic are used to describe such compounds.

Efflorescence

Some crystalline salts will lose part or all of their water of crystallization when they are exposed to the atmosphere to form a lower dehydrate or the anhydrous salt. This phenomenon is known as efflorescence and the salt is said to be efflorescent. Example is the washing soda molecule which loses nine out of its ten molecules of water of crystallization when exposed to the atmosphere.

$NaCO_3.10H_2O_{(s)} \rightarrow Na_2CO_3.H_2O(s) + 9H_2O_{(g)}$

Deliquescence

Some compounds tend to absorb a large amount of water on exposure to the atmosphere so that they eventually turn into solutions. This phenomenon is called deliquescence and substances are said to be deliquescent

Examples of these substances are

- sodium hydroxide
- iron(iii) chloride
- potassium hydroxide
- calcium chloride
- magnesium chloride
- phosphorus (v) oxide
- silica gel

A deliquescent substance such as calcium chloride or silica gel is used to maintain a dry atmosphere in a laboratory dessicator.

Hygroscopy

Hygroscopic substances also absorb moisture on exposure to the atmospere. If they are solids, they will not form solutions but merely besome sticky or moist. A hygroscopic liquid like concentrated tetraoxosulphate(vi) acid will absorb water from the air usually diluting iself 3 times its original volume. Examples of hygroscopic substances are

- sodium trioxonitrate (v)

- copper (ii) oxide
- quick lime

Drying Agents

Drying agents or dessicants are substances that have strong affinity for moisture or water. They may be either hygroscopic or deliquescent. A **drying agent** is a **chemical** used to remove water from an organic compound that is in solution. In making or isolating **chemical** compounds they often become contaminated with water. They are usually used to dry gases in the laboratory, They are also commonly used in desiccators. Example, concentrated tetraoxosulphate(vi) acid cannot be used to dry ammonia since they react to form ammonium tetraoxosulphate(vi). Commonly used drying agents are calcium chloride ($CaCl_2$), sodium sulfate (Na_2SO_4) Calcium sulfate ($CaSO_4$, also known as Drierite) and magnesium sulfate ($MgSO_4$), all in their anhydrous form.

1. Calcium chloride (n=6) is a very good drying agent for a broad variety of solvents but is generally not compatible with hydroxy (alcohol, phenol), amino (amine, amide) and carbonyl (acid, ketone, ester) functions due to basic impurities such as $Ca(OH)_2$ and CaCl(OH). In addition, it tends to form adducts with those compounds as well. Often used in drying tubes, since it also is available in granular form.

2. Calcium sulfate (n=0.5) is a neutral and good drying agent. However, it does not have a high capacity, which makes it useless for very wet solutions. The commercially available Drierite contains cobalt chloride as indicator (dry: blue, wet: pink), which can be leached out into various solvents e.g. ethanol, DMSO, DMF, ethers, etc.

3. Magnesium sulfate (n=7) is a slightly acidic drying agent. It works well in solvents like diethyl ether, but not for ethyl acetate.

4. Sodium sulfate (n=10) has a very high capacity and is mainly used for very wet solutions. It is very efficient in ethereal solutions, but it also absorbs other polar compounds like alcohols, etc.

5. Potassium hydroxide (KOH, n~1) and potassium carbonate (K_2CO_3) are both of basic nature and often used to dry basic solutions containing amines. They cannot be used to dry acidic compounds since they react with them.

6. Sulfuric acid (H_2SO_4) and phosphorous pentoxide (P_4O_{10}) are both acidic drying agents that are mainly used in desiccators and not in direct contact with the solution since they are very aggressive reagents. Both have a very high capacity. Sulfuric acid forms hydrates while phosphorous pentoxide is converted into phosphoric acid.

Assessment

1. In Universal indicators, weak alkalis are indicated through a pH of
   a. 4 to 7
   b. 8 to 10
   c. 10 to 14
   d. 7 to 14
2. Find the pOH of a chemical solution if the hydroxide ion concentration is 7.22×10^{-4} M.
3. Find the $[OH^-]$ given the pH or pOH. You are given that the pH = 8.5.
4. Find the hydroxide ion concentration of a solution with a pOH of 3.90.

Answers

1. B

Week 6

Topic: Carbon

Introduction

Carbon forms the largest number of compounds, next only to hydrogen. It ranks seventeenth in the order of abundance in the earth's crust. Carbon occurs in the free native state as well as in the combined state. Carbon and its compounds are widely distributed in nature.

In its elemental form, carbon occurs in nature as diamond and graphite. Coal, charcoal and coke are impure forms of carbon. The latter two are obtained by heating wood and coal in the absence of air, respectively. In the combined state, carbon is present as carbonate in many minerals, such as hydrocarbons in natural gas, petroleum etc. In air, carbon dioxide is present in small quantities, (0.03%).

Our food also contains carbon in the combined form. All living systems contain carbon compounds. Indeed, life as we know today, would be impossible without such carbon compounds.

Carbon is a non-metallic element and the first member of group 4 of the periodic table

Allotropes of carbon

The existence of one element in different forms, having different physical properties, but similar chemical properties is known as allotropy. Different forms of an element are called 'allotropes' or allotropic forms. Carbon shows allotropy. The various allotropic forms of carbon can be broadly classified into two classes.

- Crystalline form

Diamond and graphite are crystalline forms.

- Amorphous form

Coal, coke, charcoal (or wood charcoal), animal charcoal (or bone black), lamp black, carbon black, gas carbon and petroleum coke are amorphous forms.

Diamond

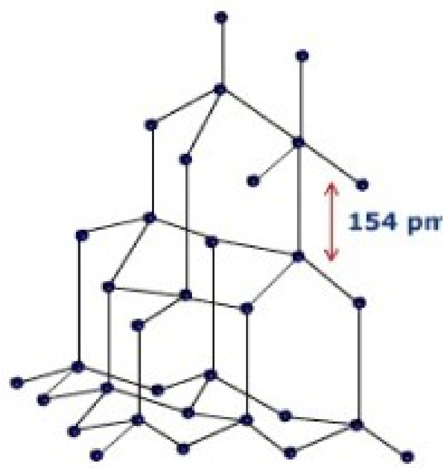
Diamonds are chiefly found in the Union of South Africa, the Belgian Congo, Brazil, British Guiana, India etc. Diamond was discovered for the first time in India. The famous 'Kohinoor diamond' (186 – carat) and the 'Regent or Pitt' (studded in Napoleon's state sword, 136.2 carat) were found near Kistna river in South India. The 'Cullinan diamond', the largest ever found weighed 3025.75 carat (about 600 g) was mined in South Africa in 1905.

Diamonds occur in the form of transparent octahedral crystals usually having curved surfaces and do not shine much in their natural form. To give them their usual brilliant shine they are cut at a proper angle so as to give rise to large total internal reflections.

Moissan (1893) prepared the first artificial diamond by heating pure sugar charcoal and iron in a graphite crucible to a temperature of about 3000°C in an electric arc furnace.

Structure of diamond

In diamond, the carbon atoms are arranged tetrahedrally (sp^3 hybridisation of C), each C atom is linked to its neighbours by four single covalent bonds. This leads to a three-dimensional network of covalent bonds.



It is due to this, that diamond is very hard, and has high melting and boiling points. In diamond, each carbon atom is bonded to the other through regular covalent bonds. The electrons thus are held tightly between the nuclei, and there are no mobile electron to conduct electricity i.e. all the valence electrons of carbon are used up in forming the covalent bonds. Hence diamond does not conduct electricity. Diamond is also denser than graphite (density: Diamond = 3.52 g cm^{-3} Graphite =2.25 g cm^{-3}) as the Diamond structure is a closely packed structure, while the layer-to-layer large distance makes graphite to have an open structure.

Properties of diamond

- Hardest substance known to man
- Brittle (not malleable)
- Insulator (non-conductor)
- Insoluble in water
- Very high melting point

Uses of diamond

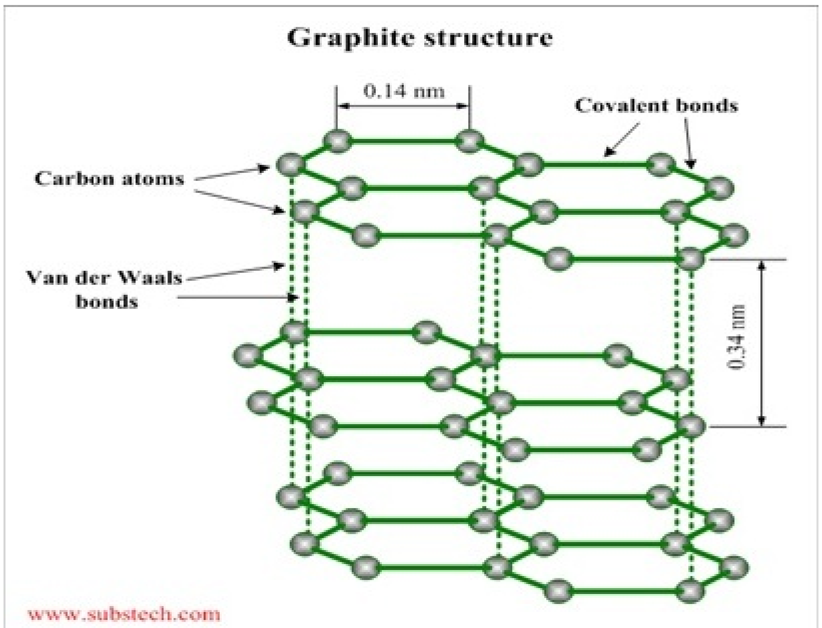
- The unusual brilliant shine of diamond makes it an invaluable precious stone in jewellery.
- Making high precision cutting tools for use in medical field.
- Because of it's hardness it is used in manufacturing tools/cutting drills for cutting glass and rock.
- In making dyes for drawing very thin wires of harder metals. Tungsten wires of thickness 1/6th that of human hair, can be drawn using diamond dyes.

Graphite

Graphite is found widely distributed in nature, viz., in Siberia, Sri Lanka, USA, Canada.

Large quantities of graphite are also manufactured from coke or anthracite in electric furnaces.

Structure of graphite



In graphite, the carbon atoms are arranged in flat parallel layers as regular hexagons. Each layer is bonded to adjacent layers by weak Van der Waals forces. This allows each layer to slide over the other easily. Due to this type of structure graphite is soft and slippery, and can act as a lubricant. Graphite is also a good conductor of electricity. In graphite, carbon atoms in each layer are bonded to three other carbon atoms by special covalent bonds. This gives some double-bond character to the C-C bonds. This gives it the presence of delocalized p-electron system. These mobile electrons explain the electrical conductivity of graphite

Properties of Graphite

- It conducts electricity
- It writes well on paper
- It has a density of 2.3g/cm^3
- It is a soft black material which feels greasy to the touch

Uses of graphite

- As a lubricant at higher temperatures.
- As a refractory material of making crucibles and electrodes for high temperature work.
- In electrotyping and in the manufacture of gramophone records: Graphite is used for making the non-conducting (generally wax) surface, so that electroplating can be done.
- For manufacturing lead pencils and stove paints.

Comparison of the properties of diamond and graphite

Diamonds and graphite are two crystalline allotropes of carbon. Diamond and graphite both are covalent crystals. But, they differ considerably in their properties.These differences in the properties of diamond and graphite are due to the differences in their structures.

Diamond	Graphite
It occurs naturally in free state	It occurs naturally and is manufactured artificially
It is the hardest natural substance known	It is soft and greasy to touch
It has high relative density (about 3.5)	Its relative density is 2.3
It is transparent and has high refractive index (2.45)	It is black in colour and opaque
It is non-conductor of heat and electricity	Graphite is a good conductor of heat and electricity
It burns in air at 900°C to give CO_2	It burns in air at 700-800°C to give CO_2
It occurs as octahedral crystals	It occurs as hexagonal crystals
It is insoluble in all solvents	It is insoluble in all ordinary solvents

Amorphous Forms of Carbon

Coal

Coal is formed in nature by the carbonization of wood. Conversion of wood to coal under the influence of high temperature, high pressure and in the absence of air is termed carbonization.

Amongst coal varieties, anthracite is the purest form. It contains about 94 – 95% of carbon. The common variety is bituminous coal; it is black, hard and burns with smoky flame.

Types of Coal

We have four types of coal namely:

- Peat – Like coal which is about 60% carbon by mass
- Lignite – Which is a brown soft coal having about 67% carbon by mass
- Bituminous – Which we use every day at home. It contains about 88% by mass of carbon
- Anthracite – This is tough and hard, having about 94% of carbon by mass

Destructive Distillation (Carbonisation) of Coal

Coal is a complex mixture of compounds composed mainly of carbon, hydrogen and oxygen with small amount of nitrogen, sulphur, phosphorus and impurities. A wide variety of substances can be obtained from it by a process known as destructive distillation of coal

The chief and important products of this process are coal gas, ammoniacal liquor, coal tar and solid coke

1. Coal Gas: This is the volatile compound which contains about 50% hydrogen, 30% methane, 10% carbon (II) oxide and small amount of other gases e.g. ethane and hydrogen sulphide
2. Ammoniacal Liquor: Ammoniacal liquor consists essentially of ammonium compounds and benzene. The ammonium compounds, can be used in the manufacture of nitrogenous fertilizers while the benzene is used for the manufacture of pharmaceutical products and as a solvent
3. Coal Tar: Coal tar is a mixture of more than 200 different substances which can be separated by fractional distillation. Most of these e.g. toluene, phenol and naphthalene are used in the synthesis of important commercial products like dyes, paints, insecticides, drugs, plastics and explosives
4. Coke: The solid residue, coke with its 98% carbon by mass burns smoothly without smoke and with high calorific value. It is used in the manufacture of gaseous fuels such as producer gas and water gas. It is also used as a reducing agent in metallurgy where it reduces metallic oxides to their respective metals e.g. in the blast furnace for the extraction of iron from its ore

Uses

Coal is mainly used:

1. As an industrial fuel in steel, power generation plants etc. It is also a domestic fuel to a limited extent.
2. For manufacture of producer gas and water gas, which are used as fuel gases.
3. For manufacturing coal tar, coke and coal gas.
4. Anthracite coal is used for preparing graphite.
5. For the manufacture of synthetic petrol by catalytic hydrogenation of coal.

Destructive Distillation of Wood

Wood is a complex substance like coal except that the percentage composition of the elements present in it is different. For example, wood has a higher percentage of carbon than coal. Destructive distillation of wood yields these four fractions:

Wood —> Wood charcoal + Pyroligneous acid + Wood tar + Wood gas

Pyroligneous acid, which in the liquid fraction contains mainly ethanoic acids, propanone, methanol and some other compounds

Charcoal

Charcoal can be made by heating wood, nut shells, bones, sugar and even blood

Wood charcoal

When wood is heated strongly in a very limited supply of air, wood charcoal is obtained. This is called destructive distillation of wood. The volatile products are allowed to escape.

Charcoal is a black, porous and brittle solid. It is a good adsorbent. Charcoal powder adsorbs colouring matter from solutions and poisonous gases from the air. Charcoal is also a good reducing agent.

Uses

- As a fuel.
- As a deodorant and in gas masks to filter pollution.
- As a discoloring agent for decolorizing oils, etc.
- In making gun powder.

Animal charcoal

Animal charcoal (or Bone charcoal) is obtained by destructive distillation of bones. It contains about 10-12% of amorphous carbon.

Sugar charcoal

It is obtained by heating sugar in the absence of air. Sugar charcoal is the purest form of amorphous carbon.

Sugar charcoal becomes activated charcoal when it is powdered to particle size of about 5 micron and heated at about 1000 K in vacuum. Activated charcoal has an increased adsorption capacity.

Lamp black

Lamp black is manufactured when tar and vegetable oils (rich in carbon) are burnt in an insufficient supply of air and the resulting soot is deposited on wet blankets hung in a room.

Lamp black is a velvety black powder. It is used in the manufacture of Indian ink, printer's ink, carbon papers, black paint and varnishes.

Carbon black

When natural gas is burnt in limited supply of air, the resulting soot is deposited on the underside of a revolving disc. This is carbon black and it is then scraped off and filled in bags.

It differs from lamp black in being not so greasy. Carbon black is added to the rubber mix used for making automobile tyres, and has replaced the use of lampblack for a number of purposes.

Gas carbon and petroleum coke

Carbon scraped from the walls of the retort used for the destructive distillation of coal is called gas carbon. During refining of crude petroleum, petroleum coke is deposited on the walls of the distillation tower.

Both, gas carbon and petroleum coke are used for making electrodes in dry cells and are good conductors of electricity.

Fuel Gases

Producer Gas: Producer gas is a mixture of nitrogen and carbon (II) oxide. This can be obtained by passing air over red hot coke in a furnace. The oxygen in the air oxidizes the coke to carbon (II) oxide. The reaction is exothermic and a large amount of heat is given off. Some carbon (IV) oxides may be formed but this is usually reduced by the hot coke to carbon (II) oxide

$2C_{(s)} + O_{2(g)} + N_2 \longrightarrow 2CO_{(g)} + N_2$

Producer gas

The mixture of the carbon (II) oxide (one third by volume) and the unchanged nitrogen from the air forms the producer gas which is tapped off as a ready gaseous fuel at the top of the furnace

Producer gas has a low heating power because it contains about 67% non-combustible nitrogen and 33% carbon (II) oxide. It is not expensive and it is widely used to heat furnaces. It is also a source of nitrogen for the manufacture of ammonia (Haber process)

Water Gas: Water gas is a mixture of equal volumes of carbon (II) oxide and hydrogen. This is produced by passing steam over coke in the furnace while it is still hot (at about 1000°C)

$C_{(s)} + H_2O_{(g)} \longrightarrow CO_{(g)} + H_{2(g)}$

Water gas

Both the hydrogen and the carbon (II) oxide in water gas burn in air releasing a lot of heat. Water gas has a high calorific value because its component gases are combustible. The demand for it is high and competitive.

Assessment

1. Which is the odd-man out?
   a. Diamond
   b. Bronze
   c. Coal
   d. Graphite

2. Ability of an element to exist in two or more physical states but maintaining the same chemical characteristics is called
   a. Isotopy
   b. Isomerism
   c. Allotropy
   d. Defraction Grating

3. When Graphite is subjected to a very high temperature and pressure for several hours in the presence of catalyst nickel, the product is
   a. soot
   b. quartz
   c. artificial diamond
   d. graphite flakes

4. The liquid product of destructive distillation of coal is
   a. ammonical liquor
   b. coal fume

c. dyestuff
d. coal plasma

5. Acheson process is the process of producing
   a. graphite from coke at high temperature
   b. producing coke using graphite at high temperature
   c. anthracite from carbon
   d. wood-charcoal from coal

Answers

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

Week 7

Topic: Oxides of Carbon

Introduction

Carbon forms two important oxides, namely carbon (iv) oxide and carbon (ii) oxide. The atmosphere contains about 0.03% by volume of carbon (iv) oxide. A small percentage of carbon (iv) oxide is found in the dissolved form of water. In the combined form, it is found mainly as metallic trioxocarbonates (iv) and hydrogen trioxocarbonates (iv) in the earth's crust.

Carbon (ii) Oxide

Carbon (II) oxide, also known as carbon monoxide, exhibits oxidation state of +2. It is the choking gas produced during the incomplete combustion of gasoline in car engines. Carbon monoxide is a colourless, odourless, and tasteless gas that is slightly less dense than air.

Laboratory Preparation of Carbon (II) Oxide

Carbon (II) oxide (carbon monoxide), CO, is prepared in the laboratory by dehydrating methanoic acid (formic acid), HCOOH or ethanedioic (oxalic) acid and passing carbon(IV) oxide, CO_2, through red-hot carbon.

Dehydration

Methanoic acid, COOH, dehydrated in the presence of a dehydrating agent, concentrated tetraoxosulphate(VI) acid, H_2SO_4.

$HCOOH(l) + \text{conc. } H_2SO_4 \rightarrow CO(g) + H_2O(g)$

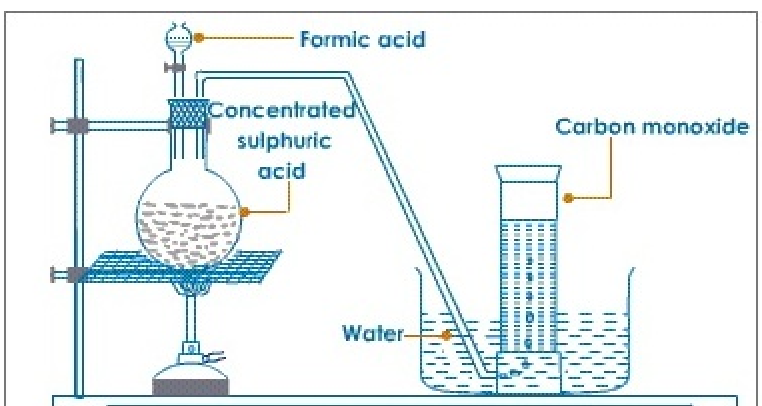


Fig. 11.3 Prepartion of carbon monoxide from formic acid

Ethanedioic acid is heated warmly with concentrated tetraoxosulphate(VI) acid to form a gaseous mixture containing equal volume of carbon(II) oxide and carbon(IV) oxide. The mixture is passed through concentrated sodium hydroxide to separate the carbon(IV) oxide, thus forming carbon(II) oxide.

$$\begin{array}{l} COOH \\ | \\ COOH(s) \end{array} \xrightarrow[-\ H_2O]{\text{conc. } H_2SO_4} CO_2(g) + CO(g)$$

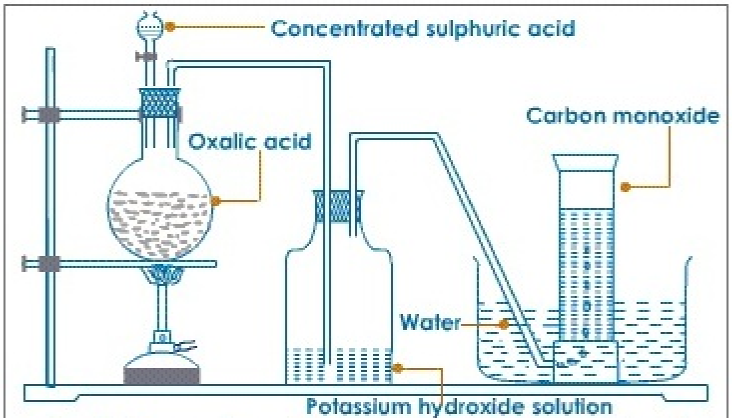


Fig. 11.2 Preparlion of carbon monoxide from oxalic acid

Large quantity of carbon(II) oxide is produced by incomplete combustion of carbon compounds like petrol. The octane, C_8H_{18}, component of petrol burns in the presence of oxygen to liberate water and form large quantity of carbon(II) oxide.

$2C_8H_{18}(l) + 17O_2(g) \rightarrow 16CO(g) + 18H_2O(l)$

Reduction

Carbon(IV) oxide is prepared in the laboratory through the reaction of marble chips with dilute hydrochloric acid. The carbon(IV) oxide is then passed through strongly heated charcoal where it is mostly reduced to carbon(II) oxide.

$CO(g) + CO_2(g) \rightarrow 2CO(g)$

But the CO gas is further passed through concentrated sodium hydroxide to remove accompanying carbon(IV) oxide impurities. Finally, the pure carbon (II) oxide is collected over water.

Physical properties of carbon monoxide

- Carbon monoxide is colourless, almost odourless and tasteless gas.
- It is very slightly lighter than air.
- Vapour Density is 14 [Vapour density of air =14.4].

- Carbon monoxide is only very slightly soluble in water. 100 volumes of water can dissolve only 3.5 volumes of the gas at S.T.P
- This is a highly poisonous gas. Air containing even less than 1% of carbon monoxide, can be fatal, if breathed in for about 10 to 15 minutes.

Chemical Properties of Carbon (II) Oxide

The chemical properties of carbon (II) oxide has to do with its combination reactions and its reducing properties as presented below:

- Combination reaction with oxygen — Carbon(II) oxide combines with oxygen by burning in air with a faint blue flame to form carbon(IV) oxide,

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

- Combination reaction with haemoglobin — Carbon(II) oxide combines with haemoglobin, Hb, in the red blood cells to form carboxyhaemoglobin, COHb.

$CO + Hb \rightarrow COHb$

This reaction is responsible for the poisonous nature of carbon(II) oxide as the haemoglobin is converted to a stable carboxyhaemoglobin that cannot transport oxygen across the body.

- Carbon(II) oxide as a reducing agent: Carbon(II) oxide is a strong reducing agent that is oxidized to carbon(IV) oxide in the presence of compounds like metallic oxides, steam and iodine(V) oxide
  It reduces lead(II) oxide, PbO, to lead metal, Pb.$PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$It reduces steam to hydrogen gas, H_2.$H_2O(g) + CO(g) \rightarrow H_2(g) + CO_2(g)$and also reduces iodine(V) oxide to iodine, I_2,$I_2O_5(s) + 5CO(g) \rightarrow I_2(s) + 5CO_2($g

Uses of carbon(II) oxide

- Carbon monoxide is used as a fuel by itself, or in the form of producer gas (mixture of carbon monoxide and nitrogen), or water gas (mixture of carbon monoxide and hydrogen). It is also present in fuel gases like coal gas.
- It is used as a reducing agent in the extraction of metals. Carbon monoxide reduces the metal oxides to metals. Usually coke is used to generate this gas. In this process coke combines with oxygen to form carbon dioxide, which gets reduced to carbon monoxide due to the lack of oxygen.
- Carbon monoxide is used in the manufacture of methyl alcohol, sodium formate, phosgene, etc.

CARBON (IV) OXIDE

Carbon (IV) Oxide, also known as Carbon Dioxide (chemical formula CO2), is a colourless, odourless gas vital to life on Earth. Carbon dioxide, CO_2, is one of two oxides of the element carbon and is the main product oxide of carbon formed from the combustion of hydrocarbon fuels. Carbon dioxide is present in the atmosphere and is produced by respiration and combustion. However it has a short time in this phase as it is consumed by plants during photosynthesis

Preparation of Carbon Dioxide

Carbon dioxide is prepared

- By treating any metallic carbonate with dilute mineral acids

CaCO3 + 2 HCl ==> CaCl2 + H2O + CO2

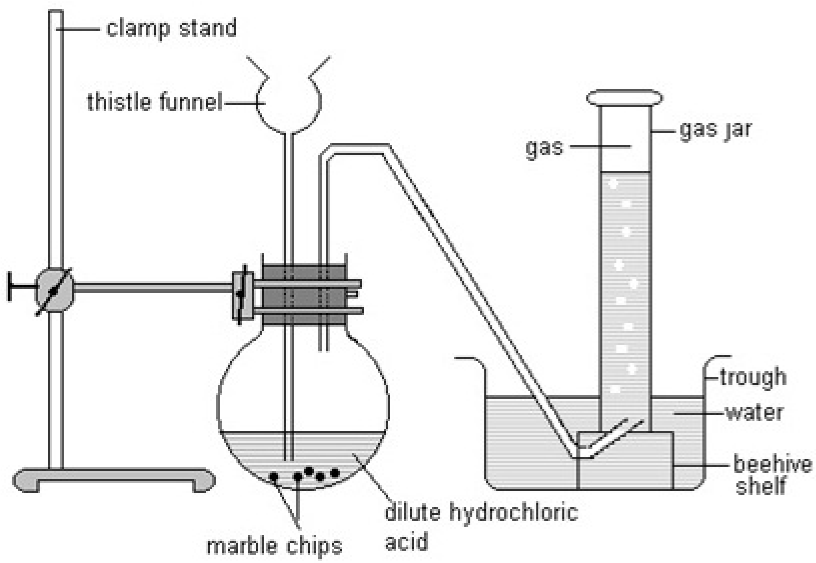


Diagram of Apparatus

- By heating carbonates of metals other than alkali metals

MgCO3 ==> MgO + CO2

Physical Properties of Carbon (IV) oxide

Some physical properties of carbon dioxide are:

- Colourless and Odourless gas with a sharp refreshing taste
- It is about 1.5 times denser than air
- Soluble in water, ethanol and acetone

- It changes blue litmus paper pink because carbon(iv)oxide dissolves in water to yield trioxocarbonate (iv) acid
- Melting point is -78°C
- Boiling Point is – 57°C

Chemical properties

- When bubbled into water it dissolves slightly and some of the carbon dioxide reacts, forming a solution of a weak acid carbonic acid which shows a pH of 5.

$H_2O(i) + CO_2(g)$ $H_2CO_3(aq)$

- It supports the combustion of only strongly burning substances such as magnesium. Magnesium metal decomposes the carbon dioxide to provide oxygen for its continued burning in the gas.

$2Mg(s) + CO_2(g)$ $2MgO\ (s) + C(s)$

This reaction is a reduction-oxidation (redox) process in that magnesium is oxidised as carbon dioxide is reduced

- When carbon dioxide is bubbled through lime water (calcium hydroxide solution), a white precipitate (calcium carbonate) is formed. This reaction is used as a test to show that a gas is carbon dioxide.

$Ca(OH)_2\ (aq) + CO_2(g)$ $CaCO_3\ (s) + H_2O\ (l)$.

If carbon dioxide is bubbled through the solution continuously then it will eventually become clear. This is because of formation of soluble calcium hydrogen carbonate solution.
$CaCO_3\ (s) + H_2O\ (l) + CO_2\ (g)$ $Ca(HCO_3)_2\ (aq)$

- Carbon dioxide reacts with strong alkalis, such as sodium hydroxide, to form carbonates. A solution of sodium hydroxide can be used to absorb carbon dioxide from the air. If excess carbon dioxide is bubbled through a solution of sodium hydroxide then a white precipitate of sodium hydrogen carbonate may be obtained.

$2NaOH\ (aq) + CO_2\ (g)$ $Na_2CO_3\ (aq) + H2O\ (l)$

$Na_2CO_3\ (aq) + H_2O\ (l) + CO_2\ (g)$ $2NaHCO_3\ (s)$

Uses of carbon dioxide

Carbon dioxide has some important uses, these are:

- To make ‘fizzy’ (carbonated) drinks like Coca-Cola, Pepsi-Cola etc. Carbon dioxide dissolves in water under pressure. Carbon dioxide gives taste to the soft drink

Soft drink

- As dry ice in refrigerators. Solid carbon dioxide sublimes and so it is used as a refrigerant for ice cream meat and soft fruits. It is used for this purpose because it is colder than ice and it sublimes at -78 °C, and so it does not pass through a potentially damaging liquid state.
- It is used in fire extinguishers; being denser than air, it blankets the fire and prevents oxygen from reaching it hence putting the fire out.

Fire extinguisher and that of it being used to put out fire

- It is produced in situ by baking powder, and also in health salts. Baking powder consists of a dry mixture of sodium hydrogen carbonate and a solid acid such as tartaric or citric acid. Reaction only takes place when water is added, when the acid reacts with the hydrogen carbonate to form carbon dioxide that makes the dough rise. A similar principle is used in health salts like Andrews Liver Salt. In health salts, carbon dioxide evolved helps in relieving indigestion or constipation
- Carbon dioxide gas is used for transferring heat in some nuclear stations in atomic reactors.

Assessment

1. Where else is CO_2 found in free state apart from the atmosphere
   a. in carbonated trees
   b. dissolved form in water
   c. in corals
   d. in limestone region

2. It is dangerous to stay in a badly ventilated room which has a charcoal fire because of the presence of
   a. carbon(II)oxide
   b. carbon(IV)oxide
   c. hydrogen sulphide
   d. producer gas

3. Carbon(iv)oxide dissolves in water to yield ....
   a. trioxocarbonate (iv) acid
   b. carbon monoxide and oxygen
   c. carbon oxosulphate
   d. None of the above

4. .... treating any metallic carbonate with dilute mineral acids or heating carbonates of metals other than alkali metals

Answers

1. B

2. A

3. A

4. Carbon Dioxide

Week 8

Topic: Trioxocarbonates

Introduction

These are the salts of carbonic acid. The anions are represented as:

- Carbonate ion trioxocarbonate (IV) : CO_3^{2-}
- Bicarbonate ion or hydrogen trioxocarbonates: HCO_3^-

These anions are formed from carbonic acid, H_2CO_3 as follows:

H_2CO_3 <——-> HCO_3^- + H^+ <——-> CO_3^{2-} + H^+

COMPOUNDS CONTAINING TRIOXOCARBONATE (IV) OR HYDROGEN TRIOXOCARBONATE (IV) ANIONS

Usually metal ions with bigger atomic size form stable trioxoxcarbonate (IV) and hydrogen trioxocarbonate (IV). Some of the hydrogen trioxocarbonate (IV) can only be detected in aqueous medium. Some important trioxocarbonate (IV) and hydrogen trioxocarbonate (IV) are listed below:

Trioxocarbonate (IV) & Hydrogen trioxocarbonate(IV)			
Group	compound	General formula	Examples
Group 1 (alkali metals)	carbonates	M_2CO_3	Li_2CO_3, Na_2CO_3, K_2CO_3 etc.,
	bicarbonates	$MHCO_3$	$LiHCO_3$, $NaHCO_3$, $KHCO_3$ etc.,
Group 2 (alkaline earth metals)	carbonates	MCO_3	$MgCO_3$, $CaCO_3$, $BaCO_3$ etc.,
	bicarbonates	$M(HCO_3)_2$	$Mg(HCO_3)_2$, $Ca(HCO_3)_2$ etc.,
p-block elements	carbonates	_	Tl_2CO_3 and $PbCO_3$
Transition elements	carbonates	_	$ZnCO_3$, $CuCO_3$, Ag_2CO_3, $FeCO_3$ etc.,.

OCCURRENCE

There are several carbonate minerals present in the nature. A few of them are listed below.

Carbonate minerals

Formula	Name of the mineral
Na_2CO_3	Soda ash or Natrite
$CaCO_3$	Lime stone or Calcite or Aragonite or Chalk
$MgCO_3$	Magnesite
$CaCO_3.MgCO_3$	Dolomite
$BaCO_3$	Witherite
$PbCO_3$	Cerrusite
$FeCO_3$	Siderite
$CuCO_3.Cu(OH)_2$	Malachite
$2[CuCO_3].Cu(OH)_2$	Azurite

Trioxocarbonates (iv)

Trioxocarbonate(iv) acid is a dibasic acid. It forms two series of salts.

- the normal trioxocarbonates (iv)
- the acidic hydrogentrioxocarbonates (iv)

Trioxocarbonates (iv) are formed naturally when trioxocarbonate (ivO acid, formed when carbon (iv) oxide dissolves in water, reacts with free metals, metallic oxides or other dissolved salts. Metallic trioxocarbonates (iv) are usually found as natural ores or deposits.

PREPARATION

Preparation of Soluble Trioxocarbonates (iv)

Sodium, Potassium and Ammonium trioxocarbonates (iv) are soluble in water. They are prepared in the laboratory by bubbling carbon(iv) oxide through a solution of the corresponding alkali.

Trioxocarbonate (IV) is formed when carbon dioxide gas is dissolved in water.

$H_2O + CO_2 \longleftrightarrow H_2CO_3$

Though looking simple, this reaction is the basic principle involved in the manufacture of club soda, coca cola, Pepsi etc. These beverages are made by dissolving carbon dioxide gas in water at high pressures. Of course, some other ingredients are also added to improve the taste of the product. That is another story. When you open the bottle, the carbon dioxide gas will come out with effervescence.

It is possible to get either trioxocarbonate (IV) or hydrogen trioxocarbonate (IV) by passing carbon dioxide into alkaline solutions.

Usually trioxocarbonate (IV) are formed when small amounts of carbon dioxide are passed through alkaline solutions.

E.g.

$2NaOH + CO_2 \longleftrightarrow Na_2CO_3 + H_2O$

small amount fairly soluble in water

$Ca(OH)_2 + CO_2 \longleftrightarrow CaCO_3\downarrow + H_2O$

small amount insoluble in water

But hydrogen trioxocarbonate (IV) are eventually formed when excess of carbon dioxide is passed into the solution.

E.g.

$NaOH + CO_2 \longleftrightarrow NaHCO_3$

excess sparingly soluble in cold water

$Ca(OH)_2 + 2CO_2 \longleftrightarrow Ca(HCO_3)_2$

excess soluble in water

Application 1: It is observed that lime water, $Ca(OH)_2$ turns milky initially when carbon dioxide is passed through it and becomes clear after passing excess of carbon dioxide. Initially an insoluble white solid,

$CaCO_3$ is formed. Hence lime water turns milky. It is then converted to water soluble bicarbonate, $Ca(HCO_3)_2$ upon passing excess of carbon dioxide by making the solution clear again.

The reactions are summarized below:

$Ca(OH)_2$ + CO_2 ———> $CaCO_3$ + H_2O

Slaked lime small amount white solid

$CaCO_3$ + H_2O + CO_2 ———> $Ca(HCO_3)_2$

excess soluble

Note: The formation of calcium trioxocarbonate (IV) is one of the reaction that occurs during setting of lime mortar, which was used in the construction of old buildings

Application 2: It has been observed that a white precipitate is formed when aqueous solution of sodium hydroxide is preserved for longer times in the containers which are not closed properly. It is because of the formation of insoluble $NaHCO_3$ when NaOH reacts with excess of carbon dioxide in air.

NaOH + CO_2 <——> $NaHCO_3$

excess sparingly soluble in cold water

GENERAL PROPERTIES

Physical state: Carbonates and bicarbonates are solids at room temperature. Trioxocarbonate (IV) of group 1 and group 2 elements are colourless. Whereas, the trioxocarbonate (IV) of transition elements may be coloured.

The polarizing power of the group 1 metal ions (M^+) is less than the polarizing power of group 2 metal ions (M^{2+}). Hence group 2 trioxocarbonate (IV) are more covalent than the trioxocarbonate (IV) of group 1.

$NaHCO_3$ and $KHCO_3$ can exist in the solid state. But the hydrogen trioxocarbonate (IV) of group 2 elements are only known in aqueous solutions.

Solubility in water: Except Li_2CO_3, The group 1 trioxocarbonate (IV) is fairly soluble in water. The solubility increases down the group as the ionic nature increases.

Group 2 trioxocarbonate (IV) are sparingly soluble in water as their lattice energies are higher (it is due to increase in covalent nature). There is no clear solubility trend observed down this group.

But group 2 trioxocarbonate (IV) are soluble in a solution of CO_2 due to formation of HCO_3^-.

Thermal stability: Carbonates are decomposed to carbon dioxide and oxide upon heating. Whereas hydrogen trioxocarbonate (IV) give trioxocarbonate (IV), water and carbon dioxide.

Thermal stability of group 1 and group 2 trioxocarbonate (IV) (also of hydrogen trioxocarbonate (IV)) increases down the group as the polarizing power of the metal ion decreases.

Due to same reason, trioxocarbonate (IV) of group 1 are more stable than those of group 2.

Small and highly charged metal ions possess more polarizing power and hence facilitates the decomposition of trioxocarbonate (IV) ion into carbon dioxide and oxide ion.

CHEMICAL REACTIONS

The most important reaction shown by these anions is 'decomposition' by liberating carbon dioxide either upon heating or by adding acids. Water or oxide are the other products.

$2HCO_3^- \longrightarrow CO_3^{2-} + CO_2 + H_2O$ (upon heating)

$HCO_3^- + H^+ \longrightarrow CO_2 + H_2O$ (in acidic medium)

$CO_3^{2-} \longrightarrow CO_2 + O^{2-}$ (upon heating)

$CO_3^{2-} + 2H^+ \longrightarrow CO_2 + H_2O$ (in acidic medium)

Illustrations:

$2NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$ (upon heating)

$NaHCO_3 + H^+ \longrightarrow Na^+ + CO_2 + H_2O$

Application: That is why baking soda ($NaHCO_3$) is used as leavening agent to raise cookies, cakes etc. It is decomposed to CO_2 and water upon heating. This makes the cookies porous and palatable.

$Ca(HCO_3)_2 \longrightarrow CaCO_3 + CO_2 + H_2O$ (upon heating)

$Mg(HCO_3)_2 \longrightarrow MgCO_3 + CO_2 + H_2O$ (upon heating)

Application: Temporary hardness of water is due to presence of bicarbonates of Ca and Mg. It is possible to remove temporary hardness by boiling water. Upon boiling, the soluble bicarbonates are decomposed to insoluble carbonates, which can be filtered off.

$CaCO_3 \longrightarrow CaO + CO_2$ (upon heating)

Application: This reaction is used to get quick lime (CaO), in lime kilns, which is further used in the preparation of slaked lime, $Ca(OH)_2$. This is also one of the reaction occurring in the manufacture of Portland cement. Technically this type of reaction is called calcination.

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

$MgCO_3 + 2HCl \longrightarrow MgCl_2 + H_2O + CO_2$

IN GENERAL

$CO_3^{2-} + 2H^+$ ———> $CO_2 + H_2O$

Calcium trioxocarbonate (IV) is present in the marble stone. This is decomposed to carbon dioxide when come into contact with acids. Hence the effervescence is observed when acids are dropped on the floor. Lime juice contains citric acid, which liberates carbon dioxide and forms insoluble calcium citrate, which appears as white marking.

Note: Effervescence is observed sometimes on granite floor which may rarely contain trioxocarbonate (IV). This may be originated from lichens lived on them.

STUDY OF SOME INDIVIDUAL TRIOXOCARBONATE (IV)

Li_2CO_3:

- Lithium carbonate is a colourless salt with polymeric nature.
- It is sparingly soluble in water and its solubility decreases with increase in temperature. But it dissolves in presence of carbon dioxide due to the formation of $LiHCO_3$.
- It is used in psychiatry to treat mania. The lithium ions interfere the sodium pump and inhibit the activity of protein kinase C (PKC).
- It is also used in the preparation of lithium cobalt oxide – which is present in lithium ion battery cathodes.

Na_2CO_3:

- Sodium carbonate is a colourless salt.
- It is fairly soluble in water.
- It is also called as washing soda.
- It is used mainly in laundries and in softening hard water.
- It is also used in making glass

Assessment

Explain the laboratory preparation of Trioxocarbonate

Week 9

Topic: Hydrocarbons

Introduction

Hydrocarbons are any organic compounds that contain only carbon and hydrogen

Hydrocarbons as a source of energy

Hydrocarbons are energy-rich because of their high carbon content. They are burned in oxygen (combustion) to release their energy.

Crude oil, or petroleum, is the main source of liquid hydrocarbons. Crude oil is a mixture of various hydrocarbons, which can be separated via fractional distillation.

These components have distinct uses, from providing energy for cooking to lubricating machine parts.

Natural gas mainly contains methane, which is the smallest hydrocarbon.

Crude oil and natural gas formed over the course of millions of years from decomposed plants and animals buried underground.

Oil, natural gas and coal (which consists of carbon and various other elements) are fossil fuels

Classification of Aliphatic Hydrocarbons

There two major classes of hydrocarbons: These are aliphatic and aromatic hydrocarbon

Aliphatic hydrocarbons made up of straight chains, branching chains, or rings that do not contain delocalized bonds

Aliphatic hydrocarbons can be divided up into two classes:

- Alicyclic hydrocarbons have carbon chains in which the two ends of the chain are joined to make a ring.
- Acyclic hydrocarbons have straight or branching carbon chains, but no rings.

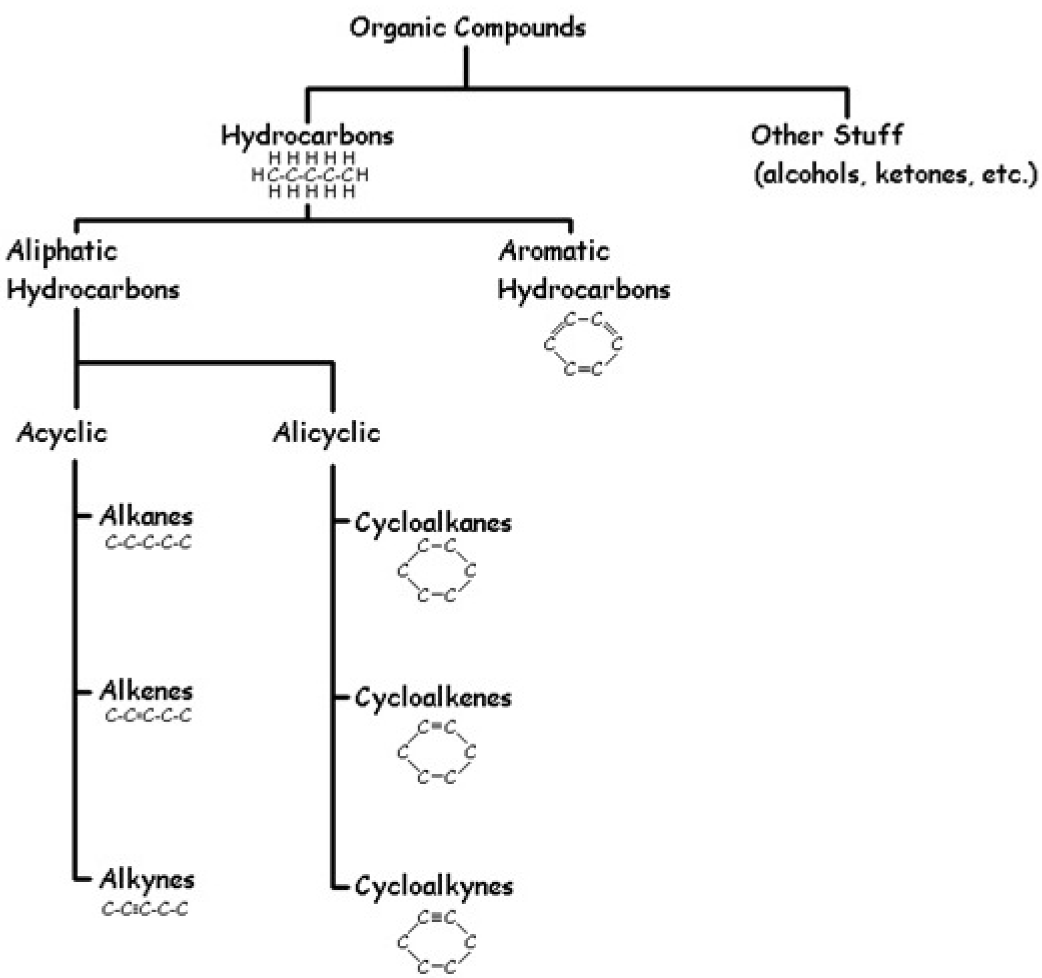
Aliphatic hydrocarbons also can be classified as to whether they contain only single bonds (saturated hydrocarbon) or if they contain more than a single bond such as double bonds or triple bonds (unsaturated hydrocarbon). Multiple bonds alter the bond geometry around the carbons. A carbon with 4 single bonds has a tetrahedral arrangement of bonds, a carbon with a double bond has a planar triangular arrangement of bonds. This changes the geometry and structure of the whole molecule.

Multiple bonds also alter the ability of the molecule to rotate around a carbon-carbon bond, that is, multiple bonds cannot rotate.

Alkanes: Contain only single carbon-carbon bonds. They have the general formula C_nH_{2n+2}

Alkenes: Contain at least one double carbon-carbon bond, but no triple bonds. They have the general formula C_nH_{2n}

Alkynes: Contain at least one triple carbon-carbon bond. They have the general formula C_nH_{2n-2}

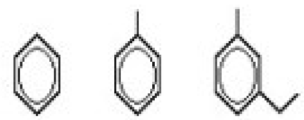


Aromatic Hydrocarbons

Aromatic hydrocarbons are distinguished by having rings with delocalized bonding orbitals.The most familiar is benzene, a 6 carbon ring: C_6H_6

Benzene rings are named just like cyclohexanes, except that the name of the ring is "benzene"

Examples:



These would be names (from left to right) benzene, methylbenzene, 1-ethyl-3-methylbenzene.

In the cases where benzene rings are attached to carbon chains that are too complicated to name as branches, the benzene ring is named as a branch instead. When benzene rings are named as branches they are called phenyl rather than benzyl.

In Summary:

Hydrocarbons

Hydrocarbons are any organic compounds that contain only carbon and hydrogen. Their general molecular formula is C_xH_y. There are several categories of hydrocarbons:

1. **Alkanes have the general formula C_nH_{2n+2}. All of the carbon-to-carbon bonds in an alkane are single bonds.**

   Example: C_3H_8 (propane)

2. **Alkenes have the general formula C_nH_{2n}. There is one carbon to-carbon double bond in an alkene. All of the other carbon-to-carbon bonds in an alkene are single bonds.**

   Example: C_3H_6 (propylene or propene)

3. **Alkynes have the general formula C_nH_{2n-2}. There is one carbon-to-carbon triple bond in an alkyne. All of the other carbon-to-carbon bonds in an alkyne are single bonds.**

   Example: C_3H_4 (propyne)

4. **Aromatic hydrocarbons are all derivatives of benzene (C_6H_6), in which the six carbons form a ring with one hydrogen bonded to each carbon.**

   Example: $C_6H_5CH_3$ (toluene)

Petroleum and Natural Gases

Introduction

Petroleum is a complex mixture of hydrocarbons and a number of different compounds whose composition varies according to its place of occurrence.

The word petroleum is composed of two Latin words Petra means rock and oleum means oil. Petroleum is a dark brown, thick and viscous liquid which occurs below the earth's surface. Petroleum is a complex mixture of hydrocarbons and different compounds of sulphur, oxygen and nitrogen in small amounts. Usually it contains alkanes, alkenes, cyclo alkanes, aromatic hydrocarbons etc.

REFINING OF PETROLEUM

The process of dividing petroleum into fractions with different boiling range volatilities and free from impurities is called refining.
The process of turning petroleum into a useful form is done in a crude oil refinery. The steps

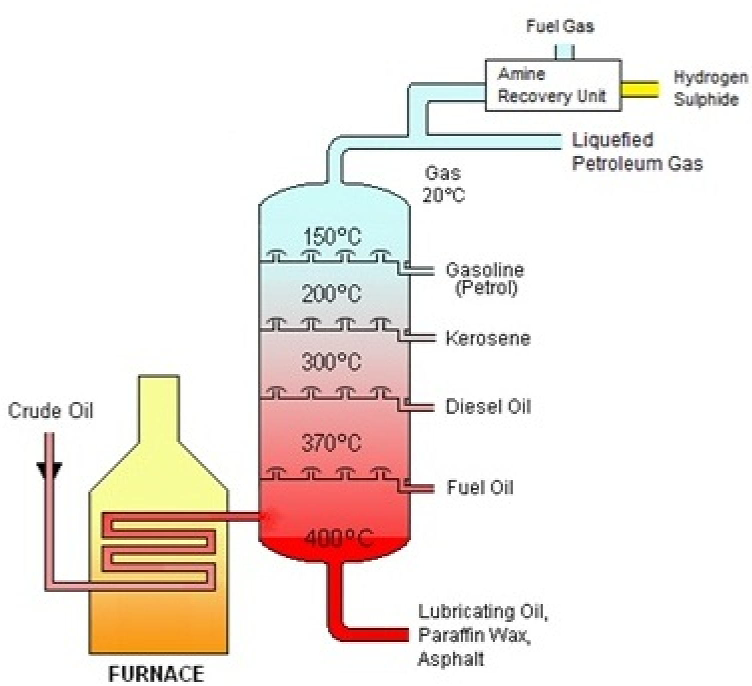
for making crude oil into oil, petrol or whatsoever are fractional distillation, cracking and reforming.

Petroleum is refined by fractional distillation. The process of separating a mixture into a series of fractions of different volatilities by means of distillation is known as fractional distillation.
In the process of fractional distillation, a mixture of different liquids is evaporated followed by condensation. Different liquids are evaporated according to their boiling point and they are collected in different chambers of distillation tower

FRACTIONAL DISTILLATION OF CRUDE OIL

The mixture is inserted at the bottom, where mostly everything will condense as the temperature is 350°C and more. The condensed crude oil will rise to the next fraction above, which has a very high temperature as well, but a slightly smaller one. Only the part of the mixture, which boiling point is under the temperature of the fraction, will condense and rise to the next fraction. The part of the mixture, whose boiling point is higher than the temperature inside the distillation fraction, will stay there and be pumped out.

Example: Supposing the temperature inside the current fraction is 300°C. The part of the mixture, which boiling point is less than 300°C, will condense. Only the molecules, which have a boiling point over 300°C, will remain in the fraction. That process is repeated about 50 to 60 times in total.



Boiling point range	Number of C- atoms	Nature
Below 20 ^{O}C	C_1 to C_4	Natural gas, bottled gas
20 ^{O}C - 60 ^{O}C	C_5 to C_6	Petroleum ether
60 ^{O}C - 120 ^{O}C	C_6 to C_7	Ligroin
40 ^{O}C - 200 ^{O}C	C_5 to C_{10}	Gasoline
175 ^{O}C - 325 ^{O}C	C_{12} to C_{18}	Kerosene oil, Jet fuel
250 ^{O}C - 400 ^{O}C	C_{12} - higher	Gas oil , fuel oil, diesel oil
Nonvolatile liquids	C_{20} - higher	Grease, lubricants
Nonvolatile solids	C_{20} - higher	Wax, asphalt, tar

Uses of petroleum fractions

The substances separated by the fractional distillation column are called fractions

Important fractions and their uses are shown in the table below:

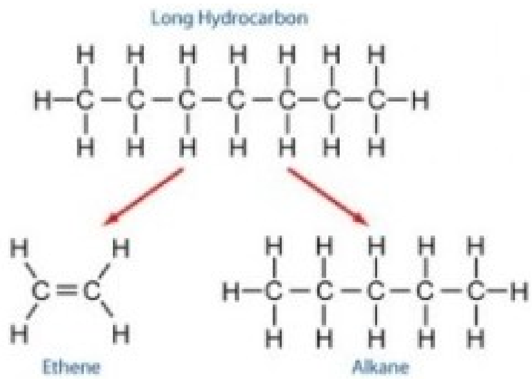
Fraction	Use	Boiling Point (∘C)
Petrol/gasoline	Fuel for petrol engines in cars	30-60
Naphtha	Industrial use, chemical feedstock	60-180
Paraffin/kerosene	Aviation fuel, cooking and heating	180-220
Diesel	Fuel for diesel engines (e.g. cars, trucks and ships)	220 – 300
Lubricating Oil	Lubricant, polishes	300-340
Bitumen	Road tar	Over 340

The fractions at the bottom (with the highest boiling points) have the lowest economic value.

For instance, bitumen (or asphalt) can only be used as road tar.

Cracking

After the fractional distillation process, the separated mixtures have to be cracked down. This means that a long molecule will be split up in smaller parts.
Firstly, single bonds will be broken down. This results into some lone electrons.
The lone electrons form double bonds. Thus, hydrogen will disconnect from the carbon atom. Hydrogen (H_2) remains as a side product. The loss of hydrogen in these smaller organic molecules is logical, because when they are lost, more lone electrons remain with what the previous lone electrons can make a bond.



Reforming

After cracking, the molecules are ready to undergo the reforming process.
This is given by the octane number. The octane number is very important in petrol. It tells what the percentage of pure heptane (in the earlier days it was octane – that is why it is called octane number) in petrol is. This is of great importance for the chemical behaviour. The quality of petrol is improved by adding mixtures to pure heptane. The chains of heptane are heated up (where platinum is used as a catalyst). So they can change. After the heating process, it shows a higher amount of branched chains. This increases the octane number.

Example: When the petrol you buy has a 98 in the name, then it means that 98% are branched chains and 2% of the mixture is pure heptane (or another pure molecule).

Octane Number

The octane number or octane rating of petrol is a measure of the proportion of branched chain hydrocarbons in a given blend of gasoline (petrol).

In other words, Octane number is a standard which determines the knocking ability and quality of gasoline. Higher is the octane number of a gasoline, lower is the knocking it produces.

Gasoline is composed of C7 – C9 hydrocarbons i.e. heptanes, octane and nonane. These hydrocarbons are present in their straight chain or branched chain isomers. It has been shown that straight chain hydrocarbons burn too rapidly in the car engine thus, causing irregular motion of the pistons which results in rattling noise. This rattling noise is known as "Knocking".

Knocking

Knocking is a sharp metallic sound produced in the internal combustion engine. Knocking is caused by the low octane number of gasoline

Synthetic Petrol

Synthetic petrol is made from materials such as coal, coke and hydrogen which do not occur in crude petroleum. Synthetic petrol can be gotten from two sources

1. From coal: When powdered coal is heated with hydrogen in the presence of iron or tin as catalyst at 500°C and 200 atmospheric pressure, it is converted into an oily mixture of hydrocarbons. The mixture is separated by distillation into a petrol fraction boiling at 200°C and a heavy oily residue which can be further treated with fresh coal to obtain more petrol

2. From Coke: When steam is passed over heated coke at 1000°C, a mixture containing equal volumes of carbon (II) oxide and hydrogen known as water gas is obtained

    1. $C_{(s)} + H_2O_{(l)} \longrightarrow CO_{(g)} + H_{2(g)}$

The water gas can be hydrogenated to a mixture of hydrocarbons by adding hydrogen and passing it over finely divided nickel as catalyst at 200°C. About half of the product is petrol, the less volatile fraction is used as fuel for diesel engines

Assessment

1. Fractional distillation involves the following process
   a. Boiling
   b. Boiling and condensation
   c. Boiling, evaporation and condensation
   d. Condensation and collection

2. Which is the odd one?
   a. Petroleum ether, petroleum gases, kerosene
   b. Gas oil, diesel lubricating oil
   c. Petroleum ether and bitumen
   d. Haemitite and asphalt

3. Which of the petrol samples are likely to cause cracking
   a. Octane
   b. 2,2,3,3 – tetra methyl butane
   c. 2,2,3 – trimethyl pentane
   d. 2,3 – dimethyl pentane

4. .... contain only single carbon-carbon bonds. They have the general formula C_nH_{2n+2}
5. .... is used to obtain petrol from the heavier or less volatile, fractions of crude oil.
6. .... contain at least one double carbon-carbon bond, but no triple bonds. They have the general formula C_nH_{2n}

Answers

1. C
2. D
3. A
4. Alkanes
5. Cracking
6. Alkenes

Week 10

Topic: Chemical Industries

The development of the chemical industry had important effects on chemistry. In 1749, the Chamber process for the commercial manufacture of hydrogentetraoxosulphate (vi) was developed.

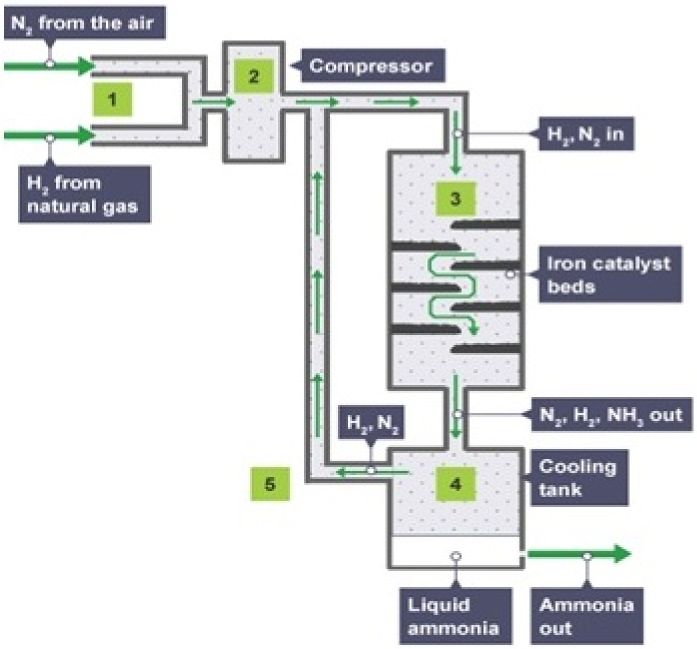
Important Raw Materials

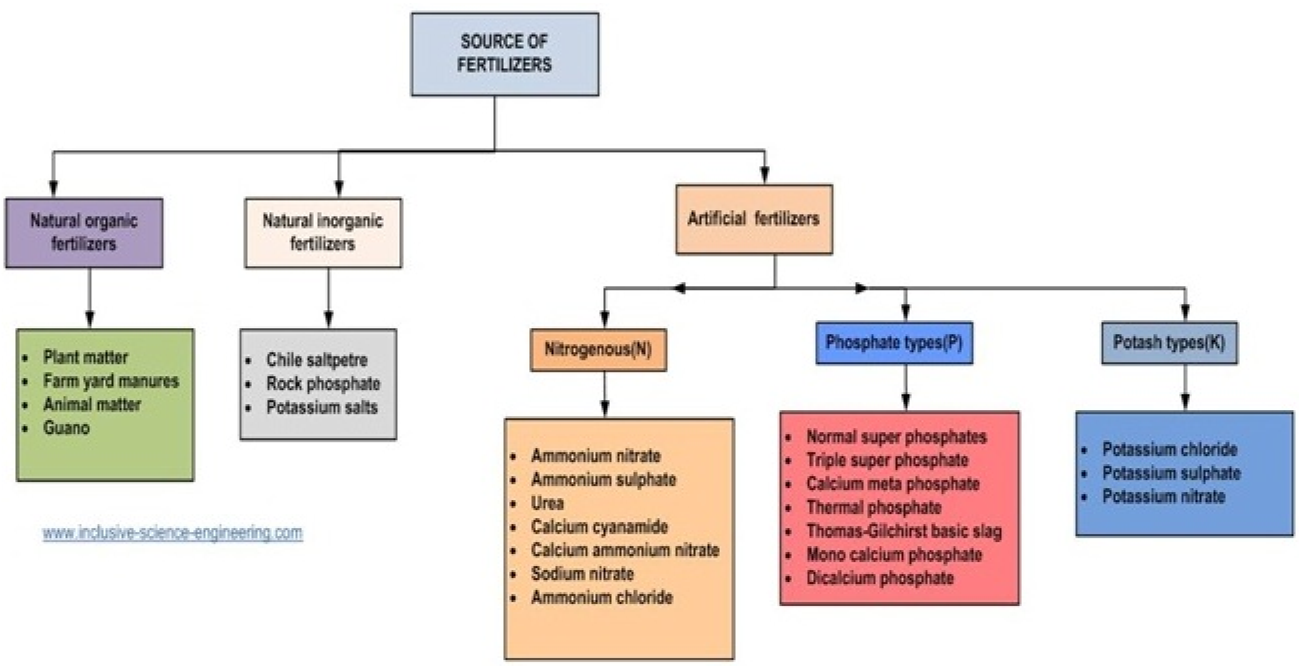
The important raw materials are air, sea-water and rock salt, sulphur, calcium trioxocarbonate (IV), metallic mineral ores, coal, natural gas and petroleum. Most of the naturally occurring raw materials are themselves chemicals which are used as starting materials to produce other chemicals and products on a large scale. All these important raw materials except air are mined from the earth's crust. Salt is obtained mainly from sea water.

Air

Air is the source of oxygen and nitrogen. Nitrogen is important because it is used in the manufacture of ammonia by haber process. Ammonia itself is used in the manufacture of trioxonitrate (V) acid (HNO_3) which is used in the making of explosives, plastics and other materials.

Ammonia can also be used in the manufacture of ammonium salts like ammonium trioxonitrate (V) NH_4NO_3 which are used in manufacturing fertilizers. The flow chart for the major chemical industries which use air as raw material is given below





Limestone, Chalk and Marble ($CaCO_3$)

Calcium trioxocarbonate (IV) $CaCO_3$ in the form of limestone, chalk and marble is the source of lime or calcium oxide which is used for making cement and concrete. It can also be the source of raw material for the production of baking powder and also for carbon (IV) oxide which can be used to produce sodium trioxocarbonate (IV) by Solvay process

Sea Water and Rock Salt

Salt or sodium chloride occurs as rock salt in underground deposits and in seawater. The sodium chloride salt is the major starting material for many chemical industries. Sea water is also the source of sodium bromide from which bromine is manufactured on a large scale.

Electrolysis of sodium chloride produces chlorine. The other main product is sodium hydroxide which is used in the manufacture of soap and textile and petroleum refining, as well as other chemicals such as sodium trioxocarbonate (IV) which is also a raw material for the glass industry and in the manufacture of detergents. It is also used in the softening of water in the public water works department

Sulphur

Sulphur is an element which is obtained naturally from the ground, or iron pyrites (FeS_2), copper pyrites ($CuFeS_2$), gypsum ($CaSO_4.2H_2O$) and anhydride which are the raw materials in chemical industries where sulphur (IV) oxide, trioxosulphate (IV) and tetraoxosulphate acids are produced. From these acids compounds like ammonium tetraoxosulphate (VI), paints, dyes, explosives fibres are manufactured

Metallic Mineral Ores (Aluminium, Iron and Copper)

Aluminium, Iron and Copper are important metals which are extracted from their respective mineral ores. The raw materials for the extraction of aluminium are aluminium silicates e.g. kaoline ($Al_2S_2O_7.2H_2O$), mica or feldspar ($K_2H_2Si_6O_{16}$), bauxite ($Al_2O_3.2H_2O$) and cryolite (Na_3AlF_3)

Coal

Industries using coal as raw material are the manufacturers of coal gas, coke and coal tar, benzene, ethyne and plastics. Coal tar is the one that contains a number of chemicals such as benzene, toluene, phenol, naphthalene and cresol

Petroleum

Petroleum is a major and a very important raw material for petro-chemicals like methane, ethene, propane, butane, ethyne, butadiene, isoprene, benzene and phenol. From these petrochemicals, other products and chemicals like fertilizer, agricultural motor fibre and plastics are made. These chemicals have their own different industries e.g. agricultural industry

The first attempt at oil exploration was in 1908 before the first world war which made it to stop. Later, petroleum products like petrol, diesel and kerosene were being marketed by foreign oil companies like Shell, Mobil, Gulf and Texaco

Heavy and Fine Chemicals

Heavy Chemicals

They are classified as heavy chemicals because they are required in large quantities in different types of chemical industries all over the world

One of the most important heavy chemicals is tetraoxosulphate (VI) acid which can be used in many chemical processes e.g. fertilizer production. Other examples of heavy chemicals are hydrochloric acid, trioxonitrate (V) acid, sodium hydroxide, slake lime and their respective salts. Metals such as Iron, copper, tin, aluminium and zinc are also heavy chemicals, as well as organic materials such as coke, coal tar, benzene, methyl benzene

Ammonia is also a heavy chemical. It is manufactured in large quantities and used to make fertilizers as well as explosives

Fine Chemicals

They are manufactured only in small quantities around the world.They form a limited class of specialized chemicals needed in certain countries and include analytical chemicals, dyes, additive, drugs and paints. The fine chemicals are produced to a very high degree of purity

Types of Industry

Due to different types of naturally occurring raw materials, many chemical industries are established. Some of these industries are:

- Fertilizer industry
- Paint industry
- Cement industry
- Plastic industry
- Pharmaceutical industry

Fertilizer Industry: The fertilizer industry is an important industry. Fertilizers such as ammonium trioxonitrate (V) (NH_4NO_3), ammonium tetraoxosulphate (VI) $(NH_4)_2SO_4$ and urea, as well as pesticides, insecticides, germicides, herbicides and fungicides are all products of petrochemical industry which helps in agricultural production

Paint Industry: Paint is a fluid mixture which contains suspended colouring materials. The main use of paints is for decoration and protection against weather and corrosion

Cement Industry: Cement is produced by heating a mixture of powdered lime (Calcium oxide) and clay. When mixed with water it can be used to fasten stones and bricks together. The mixture called mortar hardens like stone when it dries. Cement is mainly used as a component of concrete. The concrete can be poured and made into various forms when it gets dry

Plastic Industry: The plastics industry is divided into four categories: bags, household and kitchen wares, industrial plastic supplies and miscellaneous items. The industrial plastic supplies cover items such as casing for radios, cassette recorders and TV sets, as well as PVC pipes and fittings for the building industry. The starting raw materials for modern plastic industry are obtained by the fractional distillation of crude oil or petroleum

Pharmaceutical Industry: Many plant extracts are known to have some medicinal properties. Before, the chemists used to isolate the active components in the extracts and use them. Nowadays, they synthesize the compounds. Many drugs we have now have cured so many diseases considered to be incurable in the past. Drugs such as quinine for malaria and insulin for diabetes are examples. Many of our synthetic drugs, syringes, surgical and implant materials in the hospitals are manufactured from petrochemical products. Disinfectants, cosmetics, detergents and soap are also products of petrochemicals.

Assessment

1. Fine chemicals have the following characteristics except
   a. they are chemically pure
   b. they are produced by batch process
   c. they are produced in large quantity because of high applicability
   d. they are produced in small quantity because of limited applicability

2. Examples of heavy chemicals include
   a. NaOH
   b. perfumes
   c. H_2SO_4
   d. NH_3

3. Plastics are polymers whose production technique involves except
   a. high pressure
   b. low temperature
   c. high temperature
   d. setting

4. Metallurgy is a scientific process which involves the following except
   a. manufacture of alloys
   b. manufacture of both natural and artificial catalysts
   c. refining of metals
   d. grading of metals

5. These are examples of chemical industries except
   a. photosynthesis
   b. solvay process
   c. electrolysis of brine
   d. contact process

Answers

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