3. Properties and applications of some elements and compounds

At the end of this chapter, you will be competent to:

- Inquire into the characteristic properties of metals, non-metals, metalloids and noble gases
- Investigate the applications of metals, non-metals, metalloids and noble gases
- Inquire into the characteristic properties of acids, bases and salts
- Explore applications of some acids, bases and salts.

3.1 Metals, non-metals, metalloids and noble gases

You have already learnt that the various substances we use in our daily life are made up of one element or two or more elements combined together. These elements can further be classified as metallic elements, non-metallic elements, those with both metallic and non-metallic properties and noble gases. One of the ways in which the elements discovered so far have been classified on a scientific basis is the Periodic Table (Fig. 3.1).

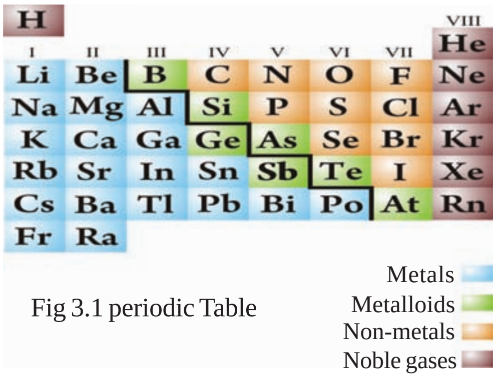

Fig 3.1 periodic Table

Observe the steps-like pattern that emerges in the elements ranging from boron in group 111 to astatine in group VII. All elements shown inside the blue cages on the left of the table are metals. Those elements next to the steps in the light green cages have properties between metals and non- metals.They are called metalloids. The elements inside the orange cages on the right side of the table are non-metals.Those elements in group viii (O) are called noble gases since they show special properties not shown by any other element.

Metals

Majority of elements (more than 80) identified so far are metals. These occur in nature in the free state as the metal itself or as compounds in combination with other elements. Metals such as gold and silver occur as the free metal, while metals such as iron, aluminium, magnesium, sodium often occur as compounds.

Physical properties of metals

Some of the physical properties of metals are, having a characteristic metallic lustre, making a sonorous sound when struck, being in the solid state at room temperature (except mercury which is a metal but is a liquid at room temperature), malleability (ability to be hammered into sheets), ductility (ability to be drawn into wires), being good conductors of heat and electricity and having high density and high melting points.

Chemical properties of metals

In general, most metals form positive ions (cations). Most metals combine with oxygen to form basic oxides. These oxides dissolve in water to form bases.

$$2Ca\ (s) + O_2\ (g) \longrightarrow 2\ CaO\ (s)$$
$$CaO(s) + H_2O(1) \longrightarrow Ca(OH)_2(aq)$$

Some metallic elements

Sodium (Na)

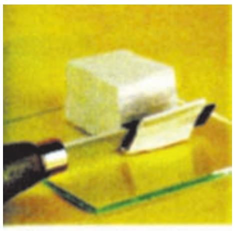

Fig. 3.2 Sodium metal can be cut easily with a knife

Among the metallic elements, sodium is one of the most reactive metals. It reacts rapidly with air or water. Therefore, sodium cannot be stored in open air. It is stored in some liquid paraffins such as kerosene oil. The reaction of sodium with air is so fast that the surface of the metal tarnishes, quickly hence the metallic lustre cannot be observed. Compared to other metals, sodium is a soft metal. It can be cut easily with a knife. The shiny metallic lustre can easily be noticed on the cut surface (Fig 3.2). Sodium atoms give a bright yellow colour to the flame. Density of sodium is 975 kgm^{-3}. Since it is lighter than water it floats on water.

Chemical properties of sodium

Sodium metal shows a strong affinity to oxygen. It reacts vigorously with oxygen in air to form the oxide.

$$4\,Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

Sodium oxide so formed reacts with water vapour in air to give sodium hydroxide. Sodium hydroxide reacts with carbon dioxide gas in air to form sodium carbonate.

$$Na_2O\,(s) + H_2O(1) \longrightarrow 2NaOH(aq)$$
$$2NaOH\,(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O\,(l)$$

Sodium reacts rapidly with cold water forming sodium hydroxide and hydrogen gas.

$$2\,Na(s) + 2H_2O\,(1) \longrightarrow 2NaOH\,(aq) + H_2\,(g)$$

Sodium reacts violently with dilute acids producing the salt of the acid and hydrogen gas (Since this is a very dangerous reaction, never attempt to do it).

$$2\,Na(s) + 2\,HC1\,(aq) \longrightarrow 2\,NaCl\,(aq) + H_2(g)$$

Uses of sodium

Sodium metal is used in,

- producing sodium cyanide needed for the extraction of gold and silver ;
- making sodium amalgam used as a reducing agent in organic chemistry and for separating the metals from compounds of titanium and zirconium ;
- producing the indigo paint used in dying blue jeans ;
- preparing drugs such as vitamin C ; and
- sodium vapour lamps.

Magnesium (Mg)

Magnesium is a very reactive, light metal. It is not found in the free state. It occurs as magnesium chloride in sea water. Its density is 1740 kg m^{-3}. He melting point is 650 °C and boiling point is 1100 °C. Magnesium slowly tarnishes on exposure to air, so the shiny nature is not easily noticed. But if cleaned with a sand paper, the shine will appear.

Chemical properties of magnesium

When heated in air magnesium burns with a bright white flame to produce magnesium oxide.

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

Magnesium does not react ~~with cold~~ water, but reacts with hot water.

$$Mg\ (s) + 2H_2O\ (1) \longrightarrow Mg(OH)_2\ (aq) + H_2\ (g)$$

Magnesium when hot reacts with steam to produce magnesium oxide and hydrogen gas.

Activity 3.1

Heat a sample of magnesium till it catches fire. Allow it to come in contact with steam as shown in the figure. Observe burning of magnesium in steam with a bright white flame.

Fig. 3.3
Reaction of magnesium with steam.

Magnesium reacts rapidly with dilute acids producing the magnesium salt and hydrogen.

$$Mg(s) + 2\ HCl(aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$$

Uses of magnesium

Magnesium is used in,

- making alloys. Magnesium is mixed with aluminium to form magnallium an alloy, which is shiny, light and resistant to corrosion. It is used in making aeroplane and vehicle parts
- making fertilizers
- making pharmaceutical drugs such as milk of magnesia.
- making fireworks and flaxes
- Preventing corrosion of iron by using as a sacrificial metal.

Fig. 3.4 Milk of Magnesia

Iron (Fe)

Iron is a commonly used, fairly reactive metal. The pure metal is soft. Hence pure iron is not very suitable for making tools and other appliances. But the strength of iron can be increased by adding other metals and making alloys. Steel is made by adding about 1.5 % of carbon to iron ; stainless steel is made by adding 12-18 % chromium and 8% nickel. A special property of iron is its attraction to magnets. Iron has a high density (7900 kg m^{-3}) when compared to metals such as sodium. Iron has a relatively high melting point (1525 °C).

Chemical properties of iron

When iron is exposed to atmospheric oxygen a reddish brown layer of oxide is formed on the surface. This is called rusting. Rust is hydrated iron oxide.

$$4\,Fe\,(s) + 3\,H_2O\,(l) + 3\,O_2\,(g) \longrightarrow 2\,Fe_2O_3\,3\,H_2O(s)$$

Iron wool, when heated in air burns easily forming iron oxide.

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

When steam is sent over red-hot iron wool or iron powder, ferrosoferric oxide and water are formed. This is a reversible reaction.

$$3\,Fe\,(s) + 4\,H_2O\,(g) \rightleftharpoons Fe_3O_4(s) + 4H_2\,(g)$$

Iron reacts with dilute acids to form the salt of the acid with the release of hydrogen.

$$Fe\,(s) + 2HCl\,(aq) \longrightarrow FeCl_2\,(aq) + H_2\,(g)$$

Uses of iron

Iron is used

- in producing alloys such as steel
- in making magnets
- to reinforce concrete
- in constructing bridges, buildings, roofing and machinery
- to make iron nails, meshes, barbed wire
- to make pharmaceutical drugs containing iron.

  Steel (an alloy of iron) is used for making tools, blades and cutlery. Stainless steel is used for making knives and blades.

Assignment 1

Prepare two lists of the metal objects in and around your home. Classify them as those made of pure iron, and those made of alloys of iron. Find out the advantages and disadvantages of using the pure metal or the alloy for each object.


Fig 3.5 Some non-metallic elements

Non-metals

Properties of non- metals are almost opposite to those of metals. Some non-metals are found in the free state but some exist as compounds in combination with other elements. Some non-metals are found as solids, e.g. sulphur, phosphorus, iodine. Bromine exists as a liquid at ordinary temperatures. Fluorine, chlorine, nitrogen, oxygen are gaseous non-metals. Non-metals are not lustrous. They are neither ductile nor malleable. Most solid non-metals are brittle. Non-metals are weak conductors of heat and electricity. Yet, graphite which is a non-metal is a good conductor of electricity. Densities of non-metals are comparatively low, yet diamond has an extremely high density.

Chemical properties of non-metals

Non-metals form negative ions (anions)

Non-metals combine with oxygen to form oxides which are acidic. These are often gases. They dissolve easily in water to form acids.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$
$$CO_2\ (g) + H_2O\ (1) \longrightarrow H_2CO_3(aq)\ \text{(Carbonic acid)}$$

Some common non-metals

Nitrogen (N)

Nitrogen exists in the atmosphere as a free gas. About 78.1% by volume of air is nitrogen. Nitrogen element is a major component in plant and animal proteins. It is found in the soil as a component in soil air as well as a component in the organic part of humus and nitrates, nitrites and ammonium compounds.

Nitrogen gas is colourless and odourless. lt is not a supporter of combustion. It is slightly lighter than air. It is slightly soluble in water. Its melting point is -195.7°C.

Chemical properties

Nitrogen is a very unreactive gas. At high temperatures, nitrogen reacts with non-metals such as oxygen, hydrogen, carbon, silicon as well as with metals such as magnesium, aluminium.

When a strong electric arc is sent through air, nitrogen combines with oxygen in air to form unstable nitric oxide. Nitric oxide so formed further reacts with oxygen to form nitrogen dioxide which is acidic.

$$N_2\ (g) + O_2(g) \longrightarrow 2NO(g)$$
$$2\ NO(g) + O_2(g) \longrightarrow 2\ NO_2(g)$$

Under special conditions, nitrogen gas reacts with hydrogen to form ammonia gas.

$$N_2(g) + 3H_2(g) \longrightarrow 2\,NH_3(g)$$

In industry, ammonia is prepared in this way. The ammonia so manufactured is used to make ammonium fertilizers and explosives.

Nitrogen gas reacts with metals such as magnesium to form the nitride of the metal.

$$3\,Mg(s) + N_2\,(g) \xrightarrow{\triangle} Mg_3\,N_2(s)$$

Uses of nitrogen

- Nitragen is used for the industrial manufacture of ammonia, chemical fertilizers and other nitrogen compounds.
- Since it is an inert gas, it is used to fill electric lamps and thermometers.
- An inert atmosphere of nitrogen is maintained in the preparation of electronic appliances.
- It is used as a blanketing gas in storing certain equipment. Packets of milk powder are filled with nitrogen gas.
- Liquid nitrogen is used as a coolant.

Sulphur (S)

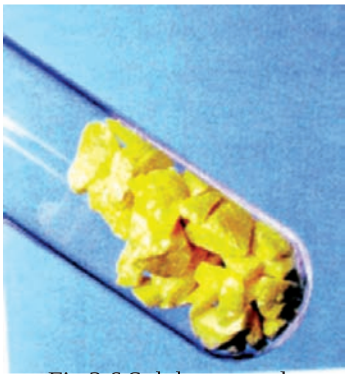

Fig 3.6 Sulphur crystals

Sulphur exists naturally in many forms. When a certain element exists in different forms in nature they are known as allotropes. Sulphur may be obtained as brittle, bright yellow crystals or as a white, non amorphous powder (Fig. 3.6). In nature it may exist in the free state or as compounds such as sulphates, and sulphides. It is a major component of amino acids in plant and animal tissues. Sulphur shows non-metalic properties. It lacks a characteristic odour or taste. It is a non- conductor of electricity. It does not dissolve in water, but may dissolve slightly in organic solvents. It is highly soluble in carbon disulphide liquid. Sulphur melts at 113°C.

Chemical properties

Sulphur burns in air with a bright blue flame to form sulphur dioxide gas.

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

On heating sulphur reacts with many metals to form the sulphide of the metal.

$$Fe(s) + S\,(s) \xrightarrow{\triangle} FeS\,(s)$$

Activity 3.2

Hold a piece of sulphur with a pair of tongs and bring it close to the flame. Observe the blue flame, and the white fumes that are given off.

Uses of sulphur

- industrial production of sulphric acid
- vulcanising rubber
- preparation of sulphites of calcium and magnesium which are used to bleach wood pulp
- preparation of sulphur containing paints
- preparation of carbon disulphide, sulphur dioxide
- making gun powder in matchstick and fireworks industry
- beer and wine industry
- The production of pharmaceuticals.
- Sulpher is also used as fungicide.

Carbon (C)

Carbon is one of the most common non-metals found freely. It occurs as carbon dioxide gas in the atmosphere. It is a component of plant and animal tissues as well as many organic compounds. It is the main component in hydrocarbons, petroleum products and coal. Carbon also shows allotropy. It exists in two crystalline forms and in amorphous form.

Crystalline forms of carbon: diamond, graphite
Amorphous forms: charcoal, coal, lamp soot

Physical properties

Physical properties of carbon differ according to the particular allotrope. Except diamond all other forms of carbon are black. They are all solids, with low densities. Yet diamond has a very high density being one of the hardest substances known. High refractive index and extreme hardness has made diamond much valuable. Carbon is a non-conductor of electricity. Yet graphite is a good conductor of electricity. Charcoal has a property of adsorbing gases.

Chemical properties

Carbon is a fairly unreactive element. At very high temperatures it combines with oxygen, but does not react with acids, bases, or chlorine. Carbon, when strongly heated, combines with oxygen to form carbon dioxide gas.

$$C(s) + O_2\,(g) \longrightarrow CO_2\,(g)$$

At high temperatures carbon reacts with calcium oxide to form calcium carbide.

$$2CaO\,(s) + 5C(s) \longrightarrow 2CaC_2\,(s) + CO_2\,(g)$$

Uses of carbon

Different forms of carbon have different applications. Listed below are some of the uses of carbon.

Form of carbon	Uses
Amorphous carbon	Pigment in black paints Vulcanizing rubber
Coal	Fuel
Graphite	Lead pencils Electrodes in cells Lubricant
Diamond	Making jewellary Cutting glass In machine parts subject to high wear and tear
Charcoal	Water purification Adsorbing gases
Carbon fibres	In spacecraft parts, space gear and missiles. These are resistant to high temperatures caused by friction.

Fig. 3.7 Carbon fibres

Chlorine

Chlorine exists as a gas at ordinary temperatures. It is a greenish yellow gas with a characteristic odour. It is heavier than air and dissolves fairly in water. Its melting point is -101°C and boiling point is - 34.5° C.

Chemical properties

Chlorine is a highly reactive gas. It reacts with many metals as well as non-metals. Chlorine has a great affinity to hydrogen. A mixture of chlorine and hydrogen gases in the presence of sunlight react explosively giving hydrogen chloride.

$$H_2(g) + Cl_2\ (g) \longrightarrow 2\ HC1\ (g)$$

Chlorine gas dissolves easily in water giving hydrochloric acid and hypochlorus acid. The latter is unstable, hence breaks down to form atomic oxygen. The atomic oxygen has a bleaching action. You can observe this bleaching action by introducing moist coloured petals of a flower to a gas jar of chlorine and observing the colour change.

The reactions bringing about bleaching are as follows:

$$Cl_2(g) + H_2O\ (1) \longrightarrow HC1\ (aq) + HOC1\ (aq)$$

$$HOCl(aq) \longrightarrow HC1\ (aq) + 'O'\ (g)$$

Coloured substance + Atomic oxygen (O) $\longrightarrow$ Bleached (colourless) substance

Uses of chlorine

Chlorine is used

- as a bleaching agent for plant material, hemp, cotton, paper, pulp, clothing etc.
- to make bleaching powder
- to sterilize drinking water and water in swimming pools
- in the extraction of gold metal
- for the production of hydrochloric acid, carbon tetrachloride and polymers such as PVC.

Metalloids

Metallods are those elements which show the properties of both metals and non-metals. Elements such as boron, silicon, germanium, arsenic, antimony belong to this group. These elements form compounds with giant molecules. They exist in the solid state at room temperature. Their melting points and boiling points are generally high. They show low electrical conductivity. But by increasing the temperature or by adding small amounts of certain impurities their electrical conductivity can be increased.

Some metalloids

Silicon (Si)

The most abundant element in the earth's crust next to oxygen is silicon. Silicon exists naturally as compounds in the crystalline form as well as in the amorphous form. The crystalline forms are quartz, sand and precious stones such as emerald. Clay is an amorphous form of silicon compounds.

Fig 3.8 Silicon sticks and pellets

Uses of silicon

- Used in making electronic equipment such as transistors diodes, and integrated circuits
- Used in making floor and glass polishes
- Many precious gems are compounds of silicon.
- Silica gel too is a silicon compound. This is used to remove moisture when packing instruments like microscopes, telescopes etc. to remove moisture.

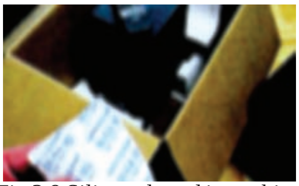

Fig 3.9 Silica gel used in packing

Germanium (Ge)

Pure germanium is a greyish white brittle solid. Its properties are much similar to those of silicon and is a semi conductor. Its melting point is 2830°C. Germanium is used in making transistors.

Boron (B)

Pure boron is a black, crystalline solid. Density of boron is 3300 kg m^{-3}. Its melting point is 2200 °C. Since its reactivity is comparatively low, hence it does not react in air even at high temperatures. Amorphous boron, at very high temperatures reacts with oxygen, nitrogen, nitric acid, concentrated sulphuric acid, carbon and sulphur to form the corresponding compounds.

Uses of boron

- making boric acid and skin ointments
- welding metals
- making certain borosilicate types of glass
- making semi- conductors
- glazing clay ornaments

Noble gases

The elements in Group VIII (Group o) of the periodic table are called noble gases. They are helium, neon, argon, krypton, xenon and radon. Since these exist in minute quantities in the atmosphere they were named rare gases. Earlier they were not known to react with any other elements and were called inert gases. Later with the discovery of compounds of xenon, krypton, and radon they are no longer considered inert. They are mostly extracted by fractional distillation of liquid air. All members are monatomic configuration. They all have a stable noble gas electron .

Helium (He)

Helium gas is present in sun's core and outer strata. This gas has got its name from the Greek word for the sun "Helios".

The amount of helium in the earth's atmosphere is as small as 0.0005%. It is colourless, odourless and non-poisonous. Its boiling point is -268.7°C and melting point is -272°C. Since it is a very light and non-combustible gas, it is often used to fill balloons. It is also used in fluorescent lamps to give coloured lights.

Neon (Ne)

Neon is a colourless, odourless, non-poisonous gas. Its melting point is -248.7 °C and boiling point is -245.90 °C.There is about 0.00123% neon in the earth's atmosphere. Neon is used in sign lamps to produce red colour.

Argon (Ar)

Earth's atmosphere contains about 0.932% of argon. Its boiling point is -186°C. Though an inert gas, argon hydrate has been reported. Argon is used for filling electric bulbs and also as a blanketing gas in metal extractions.

3.2 Acids, bases and salts

Many chemical substances are used for various purposes a in our daily lives. Among these substances, acids, bases, and salts take an important place.

Fig 3.10

Acids

Chemical compounds that produce hydrogen ions when dissolved in water, are called acids. The acidic property of an acid is due to these hydrogen ions. Hydrochloric acid in aqueous solution ionizes in the following manner:

$$HC1\ (aq) \longrightarrow H^{+}\ (aq) + CI^{-}\ (aq)$$

Strong acids and weak acids

Those acids which ionize completely in aqueous solution are strong acids while those which ionize partially are weak acids. Hydrochloric acid, sulphuric acid, nitric acid are strong acids while acetic acid, carbonic acid and phosphoric acid are weak acids.

Most acids available in the science laboratories are concentrated acids. A dilute acid solution can be prepared by adding acid to water but not vice versa.

Physical properties of acids

Acids have a characteristic sour taste. They dissolve in water. They liberate H^{+} ions in aqueous solution and conduct electricity. Acid solutions turn blue litmus red and have a pH value less than seven.

Chemical properties of acids

The more reactive metals react with acids to form the salt of the metal and hydrogen gas.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$$

Acids react with carbonates or bicarbonates to form carbon dioxide.

$$Na_2CO_3\ (s) + 2HC1\ (aq) \longrightarrow 2NaCl\ (aq) + H_2O\ (1) + CO_2\ (g)$$

$$NaHCO_3\ (s) + HC1\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (1) + CO_2\ (g)$$

Acids react with bases to form salts and water.

$$H_2SO_4\ (aq) + 2KOH\ (aq) \longrightarrow K_2SO_4(aq) + 2H_2O(l)$$

Bases

Chemical compounds that produce hydroxide ions (OH^-) when dissolved in water are called bases. The basic property of a base is due to this OH^- ion. Sodium hydroxide in aqueous solution ionizes in the following way.

$$NaOH\ (aq) \longrightarrow Na^+\ (aq) + OH^-\ (aq)$$

Metal oxides dissolve in water to form bases. There are strong bases as well as weak bases.

Those bases that ionize completely in aqueous solution are the strong bases. Sodium hydroxide, potassium hydroxide are some strong bases. Ammonium hydroxide or ammonia is a weak base since it ionizes incompletely in aqueous solution.

Among bases, those that dissolve in water are called alkalis. Potassium hydroxide, sodium hydroxide, ammonium hydroxide are some such alkalis. Hydroxides of magnesium, aluminium are slightly soluble in water.

Physical properties of bases

Bases have a characteristic bitter taste. They are slimy to touch. They conduct electricity in aqueous solution. Bases turn red litmus blue.The pH value of solutions of bases is more than seven (> 7).

Chemical properties of bases

Bases react with acids to form salts and water.

$$NaOH\ (aq) + HC1\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (1)$$

This type of reactions are called neutralisation reactions.

Applications of the acid base neutralisations

- A base such as milk of magnesia (magnesium hydroxide) is used as an antacid to neutralise the acidity in the stomach.
- A base such as wood ash or lime (calcium oxide) is used to remove the acidity of soil.
- Bee stings cause pain due to the acidic poison that enters the skin. Washing the wound with a weak base such as baking soda ($NaHCO_3$), or calcium hydroxide ($Ca(OH)_2$) will reduce the pain.
- Wasp sting is basic. Therefore, a dilute weak acid such as acetic acid or lime juice may be applied on the wound.

Salts

Salts are compounds formed by the replacement of one or more H^+ ions of an acid by a positive ion or a cation. Typically, salts are solid, ionic compounds.

$$HC1\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (1)$$

The salt formed in such a way would show acidic, basic or neutral properties depending on the strength of the acid or the base involved in the reaction.

Complete salts and partial salts

Complete salts are those formed by the replacement of all the hydrogen atoms in the acid by the metal ions.

$$H_2SO_4\ (aq) + 2NaOH\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2H_2O\ (1)$$

An acidic salt is one where only part of the hydrogens in the acid is replaced by metal ions.

$$H_2SO_4\ (aq) + NaOH\ (aq) \longrightarrow NaHSO_4\ (aq) + H_2O\ (1)$$

Here, since a hydrogen atom still remains in the salt, the salt shows acidic properties.

Hydrated salts

Certain salts have water molecules incorporated in the molecular composition. Such salts are called hydrated salts.

Hydrated copper sulphate ($CuSO_4 .5H_2O$) and hydrated sodium carbonate ($Na_2CO_3.10\,H_2O$) are examples of hydrated salts. These salts when heated become anhydrous due to removal of water molecules.

$$CuSO_4 .5H_2O\ (s) \xrightarrow{\Delta} CuSO_4\ (s) + 5H_2O\ (g)$$

Hydrated salt (blue crystals)   Anhydrous salt (white powder)

Indicators

Acidic or basic substances have the ability to change the colour of certain compounds. These substances which change colour at different ranges of acidity or basicity are called indicators. Such substances can be used to identify acids and bases. pH papers prepared by making paper to absorb an indicator (pH indicator) can also be used for this purpose. Some such indicators are litmus, methyl orange and phenolphthalein.

The colour changes shown by certain indicators in the presence of acids and bases are shown in Fig. 3.11.

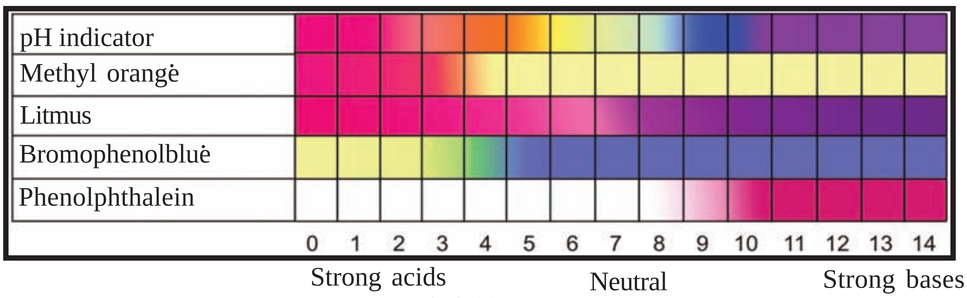

Fig 3.11

Assignment 2

Make a collection of information such as names and uses of the acids, bases and salts used in our daily life.

pH indicators

pH indicator indicates not only the acidity or basicity of a substance but also its strength. pH paper is provided with a colour code indicating the range of colours pertaining to the strength of the acid or the base.

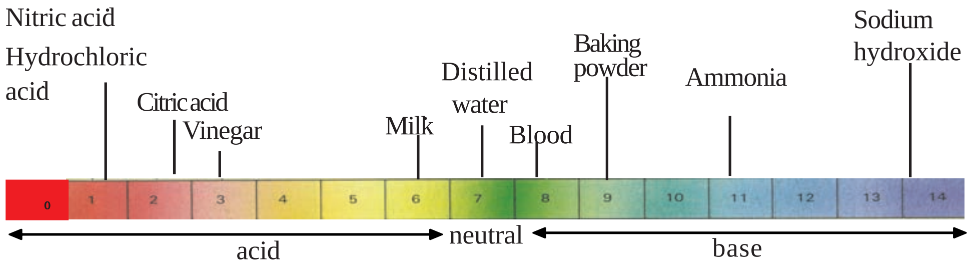

Fig. 3.12 Standard pH range

The colours indicate a range of numbers from 1 - 14. 1 shows very strong acids. Acidity decreases gradually from 1-7. From 7-14 basicity gradually increases with 14 showing strong bases.

Some commonly used acids

*Hydrochloric acid

This is a strong acid found in our laboratories. It is prepared by dissolving hydrogen chloride gas in water.

It is a colourless liquid. It dissolves in water. It has a higher density than water (1160 kg m^{-3}). It turns blue litmus red. Concentrated acid is corrosive. It will corrode materials such as paper, wood or even skin.

Hydrochloric acid forms a cloud of white fumes when in contact with ammonia gas.

Activity 3.3

Open the stopper of a conc. hydrochloric acid bottle and hold it near the mouth of a bottle of ammonium hydroxide. Observe the dense white cloud of ammonium chloride that is formed.

(conc. = concentrated)

Fig. 3.13 Reaction of hydrochloric acid and ammonia

Uses of hydrochloric acid

- removing rust from steel objects
- preparation of gelatine in the food industry
- converting starch into simple sugars
- refining common salt
- making bleaching powder for bleaching cloth, thread
- making aqua- regia, a mixture of a 3:1 conc. HC1 and conc. HNO_3 used for dissolving metals such as gold and platinum.

***Sulphuric acid**

This is the acid most commonly used in industry. It has a density higher than that of many other acids, i.e. 1840 kg m^{-3}. It is a colourless, oily liquid. Its melting point is 10.5°C. When cooled below the melting point it forms white crystals. The concentrated acid does not conduct electricity but when diluted, it is a good conductor. It has strong corrosive properties, and chars materials such as paper, wood, cloth etc. leaving only carbon. Concentrated acid has a strong affinity for water. It also can remove the elements of water from a substance, leaving only carbon. This is called dehydration.

Warning !

Concentrated sulphuric acid forms deep burns on the skin. It is dangerous to add water to concentrated acid. When preparing a sample of dilute sulphuric acid, add the acid to water gradually while stirring.

Activity 3.4

Demonstrating the dehydrating effect of sulphuric acid (Teacher demonstration) Take about three teaspoonfuls of sugar. Place in a beaker and add conc. sulphuric acid over it carefully. Observe the changes. Hydrogen and oxygen in sugar gets removed, leaving a residue of carbon.

$$C_{12}H_{22}O_{11}(s) \longrightarrow 12C\ (s) + 11\ H_2O\ (l)$$

Uses of sulphuric acid

- production of fertilizers such as ammonium sulphate, and super phosphate
- production of paints, plastics, detergents, drugs and synthetic silks
- lead accumulators as the electrolyte
- drying gases
- petroleum refineries.

*Acetic acid

It is a weak acid. It is important in industry. Pure acid is known as glacial acetic acid. It is a colourless liquid. with a melting point of 17 °C. The pure acid does not conduct electricity but dilute acid conducts to a slight degree. Dilute acid is not harmful, but the concentrated acid is corrosive. It may cause burns on the skin. Vinegar is a very dilute solution of acetic acid. Like other acids it reacts with metals, carbonates or bicarbonates, but the action is very slow.

Uses of acetic acid

- food preparation (vinegar)
- coagulating rubber from latex
- making photographic films
- paper industry
- making synthetic thread in textile industry.

Some commonly used bases

*Sodium hydroxide

Physical properties

This is a strong alkali. It's found in the form of white pellets. It has a strong attraction for water. Hence when exposed to air it absorbs moisture and becomes moist. Its melting point is 318 °C and is soluble in water. Its dissolution in water is an exothermic change. Sodium hydroxide has corrosive properties and attacks plant and animal tissues.

Chemical properties

Sodium hydroxide reacts with solutions of metal salts and forms the hydroxide of the metal.

$$CuSO_4\ (s) + 2\ NaOH\ (aq) \longrightarrow Na_2SO_4\ (aq) + Cu\ (OH)_2\ (s)$$

It also reacts with non-metals such as silicon, phosphorus, sulphur and halogens. Sodium hydroxide reacts with metals such as aluminium and zinc giving off hydrogen. Sodium hydroxide reacts with glass. Hence it should not be stored in glass vessels.

Do you know?
The reason for the soapy feeling when you touch a basic substance is because it reacts with the oil in your skin and forms a soapy subtance.

Uses of sodium hydroxide

- production of soap, paper, drugs, synthetic silks
- petroleum refining
- production of paints

*Calcium hydroxide

It is not as strong as sodium hydroxide. Calcium hydroxide is formed by adding water to quicklime. Hence it is called slaked lime.

$$CaO\ (s) + H_2O\ (1) \longrightarrow Ca(OH)_2\ (aq)$$

The colourless liquid formed by filtering a solution of calcium hydroxide in water is called lime water. If carbon dioxide is passed into lime water it turns milky, due to the formation of insoluble calcium carbonate.

$$Ca(OH)_2(aq) + CO_2\ (g) \longrightarrow CaCO_3\ (s) + H_2O\ (1)$$

If excess carbon dioxide is passed through this solution, the milkiness disappears . This is due to the reaction of calcium carbonate with carbon dioxide forming calcium hydrogencarbonate which is soluble.

$$CaCO_3\ (s) + CO_2\ (g) + H_2O\ (1) \longrightarrow Ca(HCO_3)_2\ (aq)$$

Uses of calcium hydroxide

Calcium hydroxide is used

- in the building industry to make mortar
- to make soda lime
- in laboratories for testing carbon dioxide gas.

*Magnesium hydroxide

This is a very weak base. It dissolves partially in water to form a milky suspension. In the extraction of salt from sea water magnesium salts should be removed. This is done by precipitating it as the insoluble magnesium hydroxide.

Uses of magnesium hydroxide

- Used as an antacid for stomach acidity(gastritis)
- Used to refine sugar molasses in the sugar industry

***Ammonium hydroxide**

Ammonium hydroxide is made by dissolving ammonia gas in water. It is a weak base. A solution of ammonium hydroxide gives off ammonia gas easily. Therefore a bottle containing ammonium solution can easily be identified by the strong ammonia smell.

Uses of ammonium hydroxide

- production of fertilizers
- removal of magnesium ions in sea water

Some salts in common use

Sodium chloride

Sodium chloride or common salt is very close to animal life. It is the most abundant sodium salt. Sea water consists of 3.5% of sodium chloride. It appear as white crystals at normal temperatures. It dissolves easily in water, solubility increasing slightly with increase of temperature. Its melting point is 801 °C.

Uses of sodium chloride

- preparation of food
- preserving food
- production of chlorine, hydrochloric acid, sodium hydroxide, sodium carbonate (in the Solvay process)
- for glazing clay objects
- production of soap
- in tanning of leather

***Calcium carbonate**

Calcium carbonate is a compound which exists in nature in many forms. It is insoluble in water. Calcium carbonate has crystalline forms as well as amorphous forms. Marble is a crystalline form of calcium carbonate. Chalk and limestone are amorphous forms. Calcium carbonate does not dissolve in water but with excess carbon dioxide it dissolves to form soluble hydrogencarbonate.

$$CaCO_3\ (s) + CO_2\ (g) + H_2O\ (1) \longrightarrow Ca(HCO_3)_2(aq)$$

At high temperatures calcium carbonate decomposes to give calcium oxide and carbon dioxide. In lime-kilns calcium carbonate is burnt to obtain calcium oxide (quicklime).

$$CaCO_3\ (s) \xrightarrow{\Delta} CaO\ (s) + CO_2(g)$$

Calcium carbonate reacts with acids to form the arres ponding calcium salt and carbon dioxide.

$$CaCO_3\ (s) + 2HC1\ (aq) \longrightarrow CaCl_2(aq) + H_2O\ (1) + CO_2(g)$$

Uses of calcium carbonate

- production of quicklime
- making statues and construction of houses
- making chalk

***Sodium hydrogencarbonate**

This is a basic salt found in the form of a white powder. It is soluble in water. When heated strongly (100°C) it decomposes into the carbonate and carbon dioxide.

$$2NaHCO_3(s) \xrightarrow{\Delta} Na_2CO_3\ (s) + CO_2\ (g) + H_2O\ (g)$$

Uses of sodium hydrogen carbonate

It is used;

- as baking powder. Baking powder is sodium hydrogen carbonate with a little tartaric acid.
- in fire extinguishers

***Copper sulphate**

It is a blue crystalline, hydrated salt. that dissolves easily in water. When strongly heated, it gives off water and becomes a white powder of anhydrous copper sulphate.

Uses of copper sulphate

- production of fungicides in agricultural activities
- production of chemical reagents such as Benedict's solution and Fehling's solution
- preparation of paints
- electroplating.

Exercises

1. Compare the properties of metals and non-metals.
2. Write down balanced equations for the chemical reactions given below.

   i) Sodium with water
   ii) Iron with steam
   iii) Magnesium and steam
   iv) Decomposition of calcium carbonate on strong heating

3. Give the scientific explanations for the following.

   i) Filling the empty space in powdered milk packets with nitrogen gas.
   ii) A soapy feeling on touching caustic soda.
   iii) storage of sodium metal under kerosene oil.
   iv) Formation of white fumes when an ammonia bottle stopper is held near the mouth of a bottle of conc. hydrochloric acid bottle.
   v) Formation of a black patch on a wooden table by a drop of corce. Sulpheric acid

4. Plan out an activity to demonstrate the following.

   a) Heating magnesium in steam
   b) The dehydrating property of concentrated sulphuric acid.

5. If you came across a heavy object which was buried under the soil for a long time, what activities would you conduct to find out whether it is a metallic object.