**ELEMENTS AND THEIR COMPOUDS**

HYDROGEN

Hydrogen simply means water producer this is because it burns in air to produce water. Hydrogen is a major constituent of acids bases and salts. Hydrogen was first discovered as a separate element by Henry Cavendish in 1766 and it was named as the element Hydrogen by Lavoisier.

It has three isotopic forms which are normal hydrogen, deuterium, and tritium. Tritium is radioactive and has a half-life of about twelve years. Hydrogen makes up approximately 1 percent in the atmosphere or earth’s crust. It is al so a major component of carbohydrates, fats and oils, protein etc. which are essential to all living matter.

Hydrogen on its own doesn’t really belong to any group on the periodic table

LABORATORY PREPARATION OF HYDROGEN

Hydrogen can be prepared in the laboratory by so many means which include

Action of zinc on dilute acids except HNO3

Dilute triooxonitrate five acid will not liberate hydrogen when it reacts with zinc because it’s a strong oxidizing agent .

Zn+2HCl = ZnCl2+H2

Zn+H2SO4=ZnSO4+H2

ACTION OF METALS ON WATER

Any metal above hydrogen in the activity series can liberate hydrogen when they react with water. It just depends on the activeness of the metal. For very reactive metals, such as potassium, sodium and calcium, they liberate hydrogen when they react with cold water.

For moderately active meals such as magnesium, they will liberate hydrogen when they react with hot water

For metals that are not really active such as zinc and iron, they liberate hydrogen when they react with steam.

Reaction of sodium with cold water: sodium reacts rapidly and vigorously with cold water to produce hydrogen. The reaction is very vigorous (I.e. exothermic) and should out with extreme care using a small piece of the metal

2Na+2H2O=2NaOH+H2

Lane Process: This is the reaction of hot iron on steam. Iron at red hot liberates hydrogen from steam

3Fe+4H2O= reversible Fe3O4+4H2

Iron must be at red hot before it can liberate hydrogen from steam. This reaction is called the Lane Process.

ACTION OF AMPHOTERIC METALS ON HOT CAUSTIC ALKALIS

The amphoteric metals include Aluminium, Zinc, Lead and Tin. Amphoteric metals form oxides (amphoteric oxides) which can as both acids and bases depending on the nature of what they’re reacting with.

Amphoteric oxides liberate hydrogen when they react with alkalis. However it should be noted that silicon also liberates hydrogen with alkalis

INDUSTRIAL PREPARATION OF HYDROGEN

BOSCH PROCESS

This method involves producing hydrogen from water gas (CO+H2). Water is a mixture of carbon (II) oxide and hydrogen. Carbon first reacts with water to produce water gas and thereby liberating hydrogen and if a purer and less dangerous product is needed, water is passed through the water gas using Fe2O3 or Cr2O3 as catalyst at 450 degrees Celsius thereby producing carbon (IV) oxide and hydrogen.

ACTION OF STEAM ON NATURAL GAS (METHANE)

Using nickel as catalyst, methane reacts with steam at 800 deg and 30atm to liberate hydrogen gas. This process is similar to the bosch process but methane is used instead of coke. The carbon dioxide produced in this reaction can be removed by passing it over lime water or NaOH+

BY ELECTROLYTIC METHOD

Very pure hydrogen can be produced from the electrolysis of acids, bases or/and salts.

PHYSICAL PROPERTIES OF HYDROGEN

It is colorless

Odorless

Tasteless

It is relatively light and has a density of about 14.4 times less than that of air making it the lightest known substance

It has a low boiling point of about -253

It has a low melting point

It is neutral to litmus paper

CHE1`MICAL PROPERTIES OF HYDROGEN

COMBINATION REACTION: Hydrogen has the ability to combine with a lot of elements

It burns in oxygen (air) with a pale-blue flame to produce steam. It should be noted that the reaction is highly exothermic

2H2+O2=2H2O (stedon)

Hydrogen has a strong affinity for halogens. It combines with halogens to form hydrogen halides simply said as halides

H2+Cl2=2HCl

Hydrogen reacts with metals to produce metallic hydrides also said as ionic hydrides or simply hydrides. Metallic hydrides are crystalline solids with high melting and boiling points. They conduct electricity when molten and react readily water to form hydroxides and liberate hydrogen. The reaction between metals and hydrogen should not take place in metals.

Hydrogen reacts with nitrogen to produce ammonia

Hydrogen reacts with coke to produce synthetic fiber in the process called Bregius Process

REDUCING AGENT

Hydrogen is a powerful reducing agent . It reduces oxides of metals such as copp er, lead, iron and zinc to their respective metals while itself is oxidized to water. For this reason, hydrogen is usually passed over heated copper so as to prevent the oxidation of copper by atmospheric oxygen

CuO+H2=Cu+H2O

PbO+H2=Pb+H2O

Fe2O3+3H2=2Fe+3H2O

TEST FOR HYDROGEN

Hydrogen gives a pop sound when mixed with air and releases a pale blue flame

USES OF HYDROGEN

Its largest use is in the production of ammonia

Another major use is to increase the boiling points of compounds

It is used in the hydrogenation of (unsaturated vegetable) oil to produce margarine (vanaspati fat)

It is used in the manufacture of many organic compounds such as methanol

It is (or was) used in the filling of balloons because of its lightness although it’s not advisable to fill balloons with hydrogen especially in hot weathers because hydrogen is highly combustible. Instead, helium is used because it is also light and it is not combustible.

Its liquid form can also be used as a rocket fuel

It is a constituent of gaseous fuels such as water gas (CO+H2) and coal gas (CO+H2+CH4)

It is used to produce oxy-hydrogen flame which is used by welders for cutting and/or welding metals

GROUP (VII) ELEMENTS

These are called halogens. Halogen simple means salt producer.

These elements are Fluorince, Chlorine, Bromine, Iodine and Astatine At room temperature, Fluorine is a pale yellow gas, Chlorine is a greenish yellow gas, Bromine is a dark red liquid while iodine is a purplish-black crystalline solid.

All halogens dissolve to some extent in water. Fluorine reacts vigorously and completely with water to give oxygen and hydrogen fluoride. The reactivity series such as volatility, solubility, electronegativity, oxidizing ability and tendency to enter into chemical combination decreases down the group. That means the most reactive halogen is Fluorine

Halogens are composed of covalent molecules containing two atoms (i.e. diatomic molecules) e.g. X2 where X is the halogen e.g. Cl2 F2

The general electronic configuration of halogens is … ns2np5 where n is greater than 1 and less than or equal to six.

Halogens belong to the P-block of the periodic table.

The major similarity among the halogens is the ability to displace one another from the solutions of their salts. The higher and more reactive halogen displaces the other

F2+2KCl2KF+Cl2

Cl2+2KI2KCl+I2

Br2+2KI2KBr+I2

Chlorine is the representative element and it exhibits variable oxidation states which are -1 and 5

FLUORINE

This is the most electronegative element on the periodic table. It has some unique characteristics by showing some anomalous behavior [e.g. oxidation state of -1 only and extreme reactivity]. This is partly due to its small atomic radius resulting from the absence of d orbitals in the outer shell and partly due to its high electronegativity value

CHLORINE

This is a greenish-yellow gas. It is a highly reactive halogen and therefore does not occur in nature but in combined states in compounds. It is denser than air.

Chlorine was first collected as a separate element by a Swedish scientist named Scheele in 1774

LABORATORY PREPARATION OF CHLORINE

Chlorine can be prepared in the lab by the heating of concentrated hydrochloric acid with a strong oxidizing agent such as Manganese (IV) oxide (MnO2) or Potassium tetraoxomanganate (VII) (KMnO4)

The major function of water in the diagram above is to remove the acid fumes (impurities). The concentrated H2SO4 is used to dry the gas in the third container as shown in the diagram above.

INDUSTRIAL PREPARATION OF CHLORINE

Chlorine is prepared industrially by the electrolysis of molten chlorates of sodium, potassium and magnesium

PHYSICAL PROPERTIES OF CHLORINE

It is a greenish-yellow gas

It has a choking, unpleasant and irritating smell

It is moderately soluble in water.

It is denser than air

If inhaled in small quantity, it can cause a headache but if inhaled in a large quantity, it can cause a damage to the mucus lining of the lungs

It can be easily liquefied under a pressure of about 6atm\

CHEMICAL PROPERTIES OF CHLORINE

Reaction with hydrogen: Chlorine has a strong affinity for hydrogen. It has the ability to remove hydrogen from a compound to form HCL

Oxidizing agent: Chlorine is a powerful oxidizing agent. It oxidizes the green solution of iron (II) chloride to a brown iron (III) chloride. If a metal can form more than one chloride, the higher chloride (i.e. the one in which the metal has a higher oxidation number) is formed with chlorine. For example, in the case of iron, when iron reacts with chlorine, iron (III) chloride will be formed instead of iron (II) chloride

Reaction with alkalis: The reaction of chlorine with alkalis depends on the concentration of the alkali

Dilute alkali

Cl2+2NaOH=NaClO (Sodium oxochlorate)+NaCl+H2O

Concentrated alkali:

3Cl2+6NaOH=NaClO3(sodium trioxochlorate (V))+5NaCl+3H2O

With Slaked Lime,

Cl2+Ca(OH)2=CaOCl2.H2O (Bleaching powder)

Bleaching Agent: A solution of chlorine in water is a powerful bleaching agent. The bleaching power of chlorine is due to its ability to react with water to for oxochlorate acid (HOCl)[or chlorine water]. When chlorine water is exposed to sunlight,oxygen gas will be evolved and HCl will be formed. The released oxygen combines with dye to form a colorless compound (bleached material). Hence, chlorine bleaches by oxidation

HOCl=HCl+[O]

Dye+[O] (Colored)=[Dye+O] (Colorless)

However it should be noted that clorine does not bleach dyes that contain carbon e.g. printer’s ink`

TEST FOR CHLORINE

Chlorine turns blue litmus paper red and then pink and then bleaches it

It turns starch iodide paper blue-black

USES OF CHLORINE

It is used as a powerful germicide

It is used in the sterilization of water

It is used as a bleaching agent and it is also used to produce bleaching cream

It is used to make plastics such as Polyvinyl Chloride (PVC)

It is also used to produce tetrachloromethane [CCl4] which an important solvent in dry cleaning

It is used to produce bleaching powder with the formula CaOCl2.H2O

It is used to produce AgCl a compound of chlorine which is soluble in excess ammonia and insoluble in HNO3

It is used to produce hydrogen chloride gas which is used to produce hydrochloric acid

USES OF OTHER HALOGENS

Fluorinie is used mediacally in the manufacture of toothpaste to prevent tooth decay and other oral diseases like caries

Bromine can also be used to manufacture fire extinguishers

Bromine or bromine water is used to test for unsaturation in organic compounds

Iodine is used in the treatment of goiter conditions that’s why it’s added to salt these days.

The alcoholic solution or a solution of alcohol (ethanol) and iodine is called tincture of iodine and can be used as an anesthetic

HYDROGEN CHLORIDE GAS

This gas is popularly known as marine acid gas. It can be obtained from rock salt.

This gas can be produced in the laboratory by the action of hot concentrated tetraoxosulphate (VI) acid on sodium chloride

H2SO4+2NaCl=Na2SO4+2HCl

INDUSTRIAL PREPARATION OF HCL

HCl is produced industrially by the direct combination of hydrogen and chlorine. This reaction is photo-catalytic i.e. (sun) light is the catalyst of the reaction.

The Hydrogen chloride gas later dissolves in water to form hydrochloric acid

H2+Cl22HCl

2HCl+H2O2HCl(aq)

PHYSICAL PROPERTIES OF HCL

It is colorless

It has a choking smell

It is denser than air

It is very soluble in water forming hydrochloric acid

It turns blue litmus paper red showing its acidic properties

It is very poisonous

CHEMICAL PROPERTIES OF HCL

Reaction with Ammonia: Hydrogen Chloride gas forms a dense white fume of ammonium chloride when it reacts with ammonia

HCl+NH3NH4Cl

Reaction with Silver Nitrate: Hydrogen Chloride gas forms a white precipitate of silver chloride when it reacts with silver nitrate.

HCl+AgNO3AgCl+HNO3

AgCl a compound of chlorine is soluble in excess ammonia and insoluble in HNO3 so the AgCl formed can be filtered off

AgNO3 is usually stored in a dark reagent bottle because it’s sensitive to sunlight

TEST FOR HCL

It forms a white precipitate with AgNO3

It turns blue litmus paper red

It forms a dense white fume with Ammonia gas

USES OF HCL

It is used to produce chloroethane which is an important component in plastics

It is used to remove glue from the tendon of animals

It is used in to remove oxides from metals before electroplating them. This process is called pickling.

It is used to perform the fountain experiment

It is used in the laboratory as an analytical reagent

It is used to produce hydrochloric acid

OXYGEN AND OXIDES

Oxygen makes up about 21 percent in the atmosphere. It makes up about 33 percent by volume in dissolved air. The word oxygen simply means Acid Producer.

Oxygen was first discovered by Scheele in 1772 and Priestly in 1774. It was named oxygen by Lavoisier in 1775.

The term chemistry never lived until oxygen was discovered. Oxygen is a major component of a lot of compounds such as acids bases and salts

LABORATORY PREPARATION OF OXYGEN

Oxygen can be prepared in the lab by the thermal decomposition of potassium trioxochlorate( V) using manganese (IV) oxide as catalyst

2KClO3(Mno2+Heat)2KCl+3O2

Oxygen can also be produced in the laboratory by the decomposition of hydrogen peroxide without heating. This reaction is photocatlytic

2H2O2(sunlight)2H2O+O2

Oxygen cannot be collected by the displacement of air because its density is nearly the same as that of air. It is usually collected over mercury.

INDUSTRIAL PREPARATION OF OXYGEN

Oxygen is produced on a large scale by the liquefaction of air and then the fractional distillation of the liquefied air.

TEST FOR OXYGEN

Oxygen rekindles a glowing splint since it supports combustion. However, dinitrogen (I) oxide also does the same.

Oxygen gives a brown fume gas of nitrogen (IV) oxide when it reacts with nitrogen (II) oxide while N2O doesn’t.

Also, N2O has a sweet smell while oxygen has no smell

N2O forms nitrogen with hot copper

N2O is a laughing gas because it causes uncontrollable laughter when inhaled.

PHYSICAL PROPERTIES OF OXYGEN

It is colorless

Odorless

Tasteless

Its density is almost the same as air

It has no effect on litmus paper

It supports combustion

USES OF OXYGEN

It is required for respiration and resuscitating fainting patients

It aids respiration for high mountain climbers, high altitude pilots and hospital patients having respiratory problems

Liquid oxygen known as Tonnage is a fuel used as a propellant in space rocket

It is used to remove pig impurities from pig iron

It is used to produce oxy-acetylene and oxy-hydrogen flames used by welders for welding

OXIDES OF OXYGEN

Basic Oxides: These are oxides of metals and are otherwise known as bases. Basic oxides react with acids to produce salt and water only in a process called neutralization. Most of the oxides are insoluble in water but the ones that are soluble form alkalis (hydroxides of metals) when they react with water.

E.g. Na2O, MgO, CaO,

Na2O+2HCl2NaCl+H2O

Na2O+H2O2NaOH

Acidic oxides: These are oxides of nonmetals. They are also called acid anhydrides. Acidic oxides (or acid anhydrides) dissolve in water to form acids hence the name anhydrides (not hydrated). Acidic oxides react with bases or alkalis to form salt and water only. The acid formed from these anhydrides also undergo this neutralization reaction.

CO2+2NaOHNa2CO3+H2O

CO2+H2OH2CO3

2NO2+H2OHNO2+HNO3

Amphoteric oxides: These are oxides that can behave as both acids and bases depending on where they find themselves. They react with bases (plus water) as acids and react with acids as bases to produce salt and water only. They include oxides of aluminium, zinc, lead and tin.

ZnO+2HClZnCl2+H2O

ZnO+2NaOH+h2ONa2Zn(OH)4

Neutral oxides; Thesen are oxides that are neither acidic nor basic. They are neutral to litmus paper. They include H2O, CO, N2O etc.

Peroxides: These are also metallic oxides but contain a higher amount of oxygen than ordinary oxides. They usually liberate hydrogen peroxide when in contact with a mineral acid. In peroxides, the oxidation number of oxygen is -1 instead of its normal -2. Examples of peroxides include sodium peroxide (NA2O2) potassium peroxide (K2O2), calcium peroxide (CaO2) barium peroxide (BaO2) etc

Na2O2+2HCl2NaCl+H2O2

ALLOTROPY

Allotropy is the ability

ALLOTROPES OF OXYGEN

Oxygen has only one allotrope which is ozone. Ozone could be defined as a triatomic molecule of oxygen.

Ozone can be produced when oxygen is passed through a very high electric discharge. Ozone is very reactive to remain for long in the atmosphere.

Ozone (is believed to) exist above the earth’s surface (layer) to protect the earth from receiving excess ultraviolet rays.

Ozone in its concentrated form is used as a bleaching agent. Ozone is also used to remove offensive odor from a cold storage room

The daily activities of mankind have brought about the depletion of the ozone layer. The release of chloro fluro carbons (CFCs) has brought about its depletion. This depletion has brought about global warming i.e. the increase in the average temperature of the earth. The depletion is represented as

CARBONO AND ITS COMPOUNDS

Carbon is a solid nonmetal. It belongs to group (IV) of the periodic table. Compounds of carbon are called organic compounds with the exception of oxides of carbon, carbonates. The chemistry of organic compounds is called organic chemistry. Carbon exhibits allotropy in crystalline and non-crystalline forms

ALLOTROPES OF CARBON

In the crystalline form we have diamond and graphite

In the non-crystalline (or amorphous) we have coal, coke, charcoal and soot

DIAMOND

This is a colorless transparent solid that is octahedral in shape. Diamond is the hardest natural known substance such that only diamond can cut diamond. It does not conduct electricity because it lacks mobile electrons. It has a high refractive index and can shine continuously when polished.

USES OF DIAMOND

It is used in drilling because of its hardness

It is used in making jewelry because it shines

It can also be used in merchandise or for trade

Artificial diamonds are made by subjecting graphite to a very high pressure and temperature for several hours in the presence of a catalyst such as Nickel (Ni) or Rhodium (Rh)

GRAPHITE

This is an opaque black, soft, flaky crystalline solid. It conducts electricity because it has free mobile electrons in its lattices. It is chemically inert therefore takes no part in chemical reactions. It has a density of about 2.3gcm-3

USES OF GRAPHITE

Graphite is used as a lubricant in engines where high temperature melt ordinary lubricants because of its layered structure.

It is used as electrodes in electrolytes because it conducts electricity

A mixture of graphite and clay is used as lead in lead pencils

It is used in line crubicles used for making high-grade steel

It is used for making alloys

It is used for coating iron

It is also used to lower the rate of nuclear reactions

It is applied in paint

COAL

This is gotten as a natural deposit. It is formed when plants decay in the absence of air. The major impurity in coal is sulfur

TYPES OF COAL

Peat coal-This is the simplest coal

Lignite coal: This is a brown coal

Bituminous coal: This is known as soft coal

Anthracite coal: this is the hard coal

Charcoal

USES OF COAL

The major use of coal is to supply power to steam engines

It is also used to produce important chemicals

When coal is burned, it produces oxides (compounds) of nitrogen

COKE

This is formed by heating bituminous coal at 1200C in the absence of air to drive away the volatile constituents. This process is known as the disruptive distillation of coal

Bituminous coal(1200C+heat) coke + ammoniacal liquor +coal tar + coal gas

Coke is the only non volatile residue formed

Ammoniacal liquor is a solution of ammonia in water. It is used to produce fertilizers

Coal tar is a mixture of different organic solvents which include

Coal gas is a mixture of CO, H2 and CH4

USES OF COKE

It is used as a reducing agent in the extraction of metals

It is used to produce such as water gas and producer gas

It is used to produce graphite industrially

The gasification of coke (coke + water) is used to manufacture gaseous fuels as mentioned above

Coal tar is a mixture of benzene toluene phenol and naphthalene

Coke contains 95% carbon

Coke burns smoothly without smoke

Ammonium tetraoxosulfate (VI) is used as a fertilizer [(NH4)2SO4]

CHARCOAL

This is formed by the heating of wood, sugar, animal refuse bone or even blood in a limited supply of air

Wood charcoal is the most common type of charcoal. The major impurity in wood charcoal is sulfur. The purest form of charcoal however is sugar charcoal and it can be prepared by the dehydration of sugar using H2SO4 as the dehydrating agent.

The most impure charcoal is the animal charcoal.

USES OF CHARCOAL

Generally, charcoal is used for absorbing poison or poisonous gases.

Charcoal is also used for the purification of noble gases

It is also used domestically for cooking

SOOT

This is also known as lamp black or carbon black. They are finely divided carbon atoms.

USES OF SOOT

It is used industrially to produce rubber tyres

It is also used to produce black shoe polish, carbon paper and typewriter ribbon etc.

PRODUCER GAS AND WATER GAS

Producer gas is a mixture of carbon (II) oxide and nitrogen [2CO+N2] while water gas is a mixture of carbon (II) oxide and hydrogen [CO+H2]

Producer gas is prepared by passing air over red hot coke in a furnace. This action is highly exothermic. When air is passed over heated coke, the air oxidizes coke to CO while nitrogen (in the air) remains unchanged

2C+O2+N22CO+N2+Heat

Water gas on the other hand is a mixture of CO and H2. It can be prepared by passing steam over white hot coke in a furnace of about 1000 degrees

C+H2OCO+H2

This reaction is quite endothermic

It should be noted however that both water gas and producer gas can be produced in the same plant industrially. This plant is known as producer

This process can be done by passing air and steam alternatively through the coke. The heat heated produced when producer gas is formed is sufficient for the formation of water gas.

It should also be noted that water gas is a better fuel than producer gas because in water gas, CO and H2 both support combustion while in producer gas CO supports combustion but N2 (a diluent gas) slows down combustion. Water is used to manufacture hydrogen, methanol and butanol.

GENERAL PROPERTIES OF CARBON COMPOUNDS

All the allotropes of carbon have similar properties which include

Combustion: Carbon compounds burn in (excess) air to form CO2 and in limited supply of air to form CO

C+O2CO2

2C+O22CO

The temperatures at which they burn is different though.

Reactions in Charcoal pot

Combination Reaction: At very high temperature, carbon combines with elements such as hydrogen, sulfur, aluminium and calcium.

C+2H2CH4

C+2SCS2

3C+4AlAl4C3 (Aluminium Carbide)

2C+CaCaC2

Reducing Agent: Carbon reduces oxides of less active metals to their metals at high temperature and is oxidized to CO or CO2

Fe2O3+3C2Fe+3CO

2CuO+C2Cu+CO2

It also reduces steam to hydrogen and CO

H2O+CCO+H2

Reaction with strong oxidizing agents:

COMPOUNDS OF OXIDIZING AGENT

When carbon is heated in air, it forms two oxides which are CO2 and CO

Carbon (IV) oxide

This is a colorless gas. It was first discovered by Van Helmont towards the end of the sixteenth century

It is about 0.03% in the atmosphere

LABORATORY PREPARATION OF CO2

CO2 is prepared in the laboratory by the action of dilute mineral acid on insoluble calcium carbonate CaCO3. However it should be noted that dilute H2SO4 is the only acid that cannot be used to produce CO2 because of the formation of insoluble calcium sulfate (CaSO4) which will stop the action of the acid and the carbonate.

2HCl+CaCO3CaCl2+H2O+CO2

2HNO3+CaCO3Ca(NO3)2+H2O+CO2

If CaCO3 is not used, a hydrogen trioxocarbonate (IV) (e.g. sodium trioxocarbonate (IV) [NaHCO3]) can be used.

DIAGRAM

HCl+CaCO3KHCO3CaCl2CO2

The major use of calcium chloride is to dry the gas

INDUSTRIAL PREPARATION OF CO2

Complete combustion of hydrocarbons

CH4+2O2CO2+2H2O

As a by-product of the fermentation of sugar (glucose)

C6H12O62C2H5OH+2CO2

Reduction of metal oxides by carbon or CO

2PbO+C2Pb+CO2

Fe2O3+3CO2Fe+3CO2

PHYSICAL PROPERTIES OF CO2

It is a colorless gas

It is odorless

It has a sharp refreshing taste (when mixed with water) (carbonated water)

It is slightly soluble in water

It is denser than air. It has a density of about 1.95gdm-3

It is a weak acid

It turns (moist) blue litmus paper red (or pink)

On cooling, it readily liquefies and solidifies at -78C to form solid CO2 (a white solid) known as dry ice. It can also be obtained by covering ice with blotting paper

CHEMICAL PROPERTIES OF CO2

Reaction with Alkalis (as an acid): CO2 reacts with alkaline solutions to form salt and water only in the process called neutralization

CO2+NaOHNa2CO3+H2O

In the presence of excess CO2, the carbonates change to hydrogen carbonates

Na2CO3+H2O+CO22NaHCO3

Reaction with magnesium (Burning magnesium): CO2 does not burn neither does not burn neither does it support combustion but when magnesium burns in a gas jar of CO2 to produce a carbon deposit and a white magnesium oxide ash

CO2+2Mg2MgO+C

Reaction with water

CO2+H2OH2CO3

TEST FOR CO2

CO2 turns lime water (calcium hydroxide) milky due to the formation of insoluble calcium carbonate

Ca(OH)2+CO2CaCO3+H2O

If the CO2 is in excess, the milky color disappears as a result of the formation of hydrogen carbonate

CaCO3+H2O+CO2Ca(HCO3)2

If heat is applied, the milky color reappears again

Ca(HCO3)2CaCO3+H2O+CO2

USES OF CO2

A solution of CO2 and water has a refreshing taste and can therefore be used for making soft drinks to give these carbonated drinks a refreshing taste

It is used as a fire extinguisher because it does not support combustion

Dry ice is used as a refrigerant (for perishable goods)because it sublimes and it gives a lower temperature than normal ice

It is also used as a coolant in nuclear reactions

It is used by plants for photosynthesis

It is used in the solvay process to produce NaHCO3 and Na2CO3

CARBON (II) OXIDE (CO)

This is produced naturally by the incomplete combustion of petroleum using octane (C8H18)

2C6H18+17O216CO+18H2O

LABORATORY PREPARATION OF CO

CO can be prepared in the laboratory by the dehydration of methanoic acid (also known as formic acid and can be gotten from ants) or ethanoic acid (also known as oxalic acidwhich can be gotten from vinegar) using H2SO4 as the dehydrating agent

COOH (Conc. H2SO4 (-H2O))CO

H

Methanoic

COOHConc. H2SO4-H2OCO2+CO

COOH

CO is considered as the most dangerous gas (when inhaled) because it combines with the hemoglobin to form a stable compound (carboxyl hemoglobin) and prevents the hemoglobin from transporting oxygen to various parts of the body. CO is also called the silent killer because it is colorless, odorless and tasteless.

PHYSICAL PROPERTIES OF CO

It is colorless, odorless and tasteless

It is less dense than air

It is a neutral gas and therefore neutral to litmus paper

It is relatively insoluble in water but soluble in Ammoniacal copper (I) chloride

CHEMICAL PROPERTIES OF CO

Reducing agent: It reduces lead (II) oxide, iron (III) oxide, copper (II) oxide, iodine (V) oxide and steam while itself is oxidized to CO2

PbO+COPb+CO2

Fe2O3+3CO2Fe+3CO2

CuO+CO

I2O5+5COi2+5CO2

H2O+COH2+CO2

Reaction with oxygen: CO burns in oxygen to produce CO2

2CO+O22CO2

TEST FOR CO

CO burns in air with a (pale) blue flame

USES OF CO

It is used as a reducing agent in the extraction of metals from their ores

It is used to produce gaseous fuel

It is used in the commercial production of methanol

CO+2H2  (zinc chromate at 450C, 200atm) CH3OH

Carbon cycle

This is the cycle that shows the series of ways CO2 can be released into the atmosphere or removed from the atmosphere.

DECOMPOSITIONS OF CARBONATES

K K2CO3K2O+CO2

Na

Ca CaCO3CaO+CO2

Mg

Al

Zn

Fe

Sn

Pb

Cu

Hg 2Hg2CO34Hg+2CO2+O2

Ag

Au

DIFFERENCES BETWEEN CO AND CO2

|  |  |
| --- | --- |
| CO | CO2 |
| Neutral Oxide | Acidic Oxide |
| Reducing agent | Oxidizing agent |
| Less Dense than air | Denser than air |
| Poisonous gas | Refreshing gas |

Poisonous gases such as CO, H2S, NH3, Cl2 and SO2 can be prepared in a fume cupboard in the laboratory

SULFUR

This is a yellow solid that does not conduct electricity. It belongs to group (VI) of the periodic table. It is normally found as mineral deposits in places like America.

Sulfur has been known for a long time for its medicinal and germicidal effect on the skin.\

It can be extracted from the ground by the frasch process. It exhibits allotropy.

ALLOTROPES OF SULFUR

Sulfur has both crystalline and non-crystalline allotropic forms.

In its crystalline form, we have rhombic sulfur and monoclinic or prismatic sulfur

Non-crystalline sulfur forms we have amorphous sulfur and plastic or roll sulfur.

Rhombic sulfur has a density of 2.06gcm-3

Rhombic sulfur has an S8 structured lattice

It is the only allotrope of sulfur that is stable below 96C

It has a boiling point of about 116C

Monoclinic sulfur has a density of 1.96gcm-3

It also has an S8 structured lattice too

It has a needle like shape

Monoclinic sulfur is the only allotrope of sulfur that is stable between 96 and 119C

It has a boiling point of 119C

It should be noted that the temperature 96C that is common to the two crystalline allotropes is called the transition temperature

Sulfide ores can be concentrated by the floatation.

PHYSICAL PROPERTIES OF SULFUR

It is a yellow solid

It is a bad conductor of electricity

Its density depends on its allotrope

It is insoluble in water but soluble in benzene and carbon (IV) sulfide

USES OF SULFUR

The most importance use of sulfur is to produce SO2 (for the manufacture of H2SO4)

Sulfur is used as a germicide and fungicide

Sulfur and its products(like SO2) are used as a bleaching agent

It is used in the funginization of rubber (i.e. the process of treating rubber with sulfur so as to improve its strength, toughness and durability)

COMPOUNDS OF SULFUR

Hydrogen Sulfide (H2S)

This gas smells like rotten egg

LABORATORY PREPARATION OF H2S

H2S can be produced both in the laboratory and in an industry by the action of a dilute acid on iron (II) sulfide

2HCl+FeSFeCl2+H2S

The popular drying agent conc. H2SO4 is never used to dry H2S because it will oxidize H2S to storeform

Kipp’s Apparatus: H2S can be prepared using the Kips apparatus. This apparatus is used for an intermediate supply of any gas by the action of a liquid on a solid without heating.

PHYSICAL PRPERTIES OF H2S

It is a colorless with a repulsive smell like that of rotten egg.

It is moderately soluble in water

It is denser than air

It is an acidic gas

CHEMICAL PROPERTIES OF H2S

Reaction as an acid: H2S is a weak dibasic acid that reacts with alkalis to produce salt and water only

As a precipitating agent: H2S is used to precipitate insoluble sulfides from the solution of their salts. This method is also a way of producing acids.

H2S+(CH3COO)2Pb (Lead ethanoate)PbS (Black)+2CH2COOH

H2S+ZnSO4ZnS+H2SO4

As a reducing agent: H2S is a powerful reducing agent. It reduces powerful oxidizing agent by changing their colors and then depositing sulfur. It changes KMnO4 (Mn 7+) from purple to colorless MnO2 (Mn 4+) reduction. It changes K2Cr2O7 (Cr 6+) from orange to green and then deposits sufur in both reactions

KMnO4+H2SO4+5H2SK2SO4+2MnO2+^H2O+5S

PurpleColorless

TEST H2S

H2S is the only gas that smells like rotten egg

H2S turns lead ethanoate paper to black

USES OF H2S

It is used in the analysis of ore

It can be used to separate a group of metals from one another

SULFUR (IV) OXIDE (SO2)

It is a colorless gas that smells like matches

It is produced naturally by striking matches. It burns in a pale blue flame

S+O2SO2

LABORATORY PREPARATION OF SO2

SO2 is produced in the laboratory by the action of HCl and NaSO3

2HCl+Na3SO32NaCl+H2O+SO2

PHYSICAL PROPERTIES OF SO2

It is a colorless gas with a choking smell (or pungent smell)

It smells like burning matches

It is a poisonous gas

It is very soluble in water

It is denser than air

It turns blue litmus paper red (i.e. it is acidic)

CHEMICAL PROPERTIES OF SO2

As an acid: SO2 is an acidic oxide that reacts with bases (alkalis) to produce salt and water only

SO2+2NaOHNa2SO3+H2O

Reducing agent: SO2 is a powerful reducing agent. Its reducing properties is similar to that of H2S. They both change the colors of KMnO4 and K2Cr2O7. The only difference is that SO2 does not deposit sulfur while H2S deposits sulfur

SO2+H2O+KMnO4K2SO4+MnSO4+H2SO4

SO2+K2Cr2O7K2SO4+Cr2(SO4)3+H2O

Oxidizing agent: SO2 behaves as an oxidizing agent when it comes in contact with a stronger reducing agent such as H2S

SO2+H2S2H2)+3S

Bleaching agent: A solution of SO2 in water is a powerful bleaching agent. The bleaching action of SO2 is similar to that of chlorine gas because water must be present (i.e. they must dissolve in wter) for the bleaching effect of both substances to occur. Chlorine bleaches by oxidation while SO2 bleaches by reduction. It should be however a stronger bleaching agent and its bleaching effect lasts longer than that of SO2

TEST OF SO2

SO2 turns acidified KMnO4 from purple to colorless without depositing sulfur

It smells like matches

SO2 also turns K2Cr2O7 from orange to green to colorless without depositing sulfur

USES OF SO2

The most important use of SO2 is for the production of H2SO4

It is used as a bleaching agent for materials that can be easily damaged by chlorine (e.g. rubber and silk fabric etc.)

SO2 and some of its products are used as fumigants for destroying termites and bed-bugs

It is used as a preservative in orange juice

HYDROGEN TETRAOXOSULFATE (VI) ACID (H2SO4)

This is known as the king of chemicals

It is a heavy chemical because it is produced (and also consumed) in a large quantity

Every industrial process and laboratory activity makes use of this acid directly or indirectly

LABORATORY PREPARATION OF H2SO4

H2SO4 can be prepared in the laboratory by the lead chamber process. The following reactions are done in lead chamber process

S+O2SO2

2SO2+O22SO3

SO3+H2OH2SO4

The reactions are done in a lead chamber because the reaction is highly exothermic

INDUSTRIAL PREPARATION

H2SO4 can be produced on a large scale by the contact process. The following reactions take place during the contact process for the production of H2SO4

S+O2SO2

2SO2+O2 (reversible and V2O5 catalyst)2SO3

SO3+H2SO4H2S2O7 (Oleum)

H2S2O7+H2O2H2SO4 (98% pure)

The catalyst employed in the cintact process in vanadium (V) oxide

SO2 is passed through an electric chamber to remove dusr and impurity that might affect the catalyst

SO3 is not dissolve directly in water to produce H2SO4 because the reaction is highly exothermic and this can cause the acid solution to boil

PHYSICAL PROPERTIES OF H2SO4

It is a colorless viscous liquid

It turns blue litmus paper red

It has strong affinity for water and liberates a large amount of heat when dissolved in water

Concentrated H2SO4 is highly corrosive (i.e. it can cause severe burns with the skin)

CHEMICAL PROPERTIES OF H2SO4

Oxidizing Agent: Hot Conc. H2SO4 is a powerful oxidizing agent. It oxidizes most metals and non-metals to their highest oxides while itself is reduced to SO2 and water

C+2H2SO4CO2+2H2O+2SO2

Dehydrating Agent: Conc. H2SO4 can be used as a drying and dehydrating agent. Drying is the physical removal of water. Dehydration is the chemical removal of water. Conc. H2SO4 behaves as a dehydrating agent when it gets in contact with organic compounds. The dehydration property of H2SO4 is shown below

C12H22O11Conc. H2SO4 -11H2O12C

C2H5OH conc. H2SO4 –H2OC2H4 (ethane)

COOHCO2+CO

COOH

GENERAL USES OF H2SO4

It is used to manufacture fertilizers

It is used as an electrolyte in lead-acid accumulator

It is used in refining petroleum

It is used to produce pigment used in paint

It is used in the laboratory as a typical acid, catalyst, oxidizing agent, drying agent, dehydrating agent, analytical agent, acidifying agent etc.

SO4^2- ions are tested using barium chloride

Group (VI) elements have a general electronic configuration of ns2np4 2 where n>1

Group VI elements are oxygen sulfur selenium tellurium and polonium

NITROGEN AND ITS COMPOUNDS

Nitrogen makes up about 78% of the volume of the atmosphere. It is a diluent gas because it slows down the combustion and oxidation of most metals by diluting oxygen. It was discovered by David Rutherford in 1772. It is also a constituent of all plant and animal proteins.

Certain metals such as calcium, magnesium and aluminium when burned in air combine with oxygen and nitrogen to form oxides and nitrides respectively

LABORATORY PREPARATION OF NITROGEN

Nitrogen can be prepared in the laboratory by so many methods

Nitrogen can be produced by the thermal decomposition of ammonium nitrite (NH4NO2)

NH4NO2 (heat)N2+2H2O

It can also be prepared from the atmosphere

Air [caustic soda (NaOH)] O2+N2

N2+O2(2Cu)N2 (99% pure)

From ammonium heptaoxo dichromate (VI)

(NH2Cr2O7N2+Cr2O3+4H2O

It can also be produced from the oxidation of ammonia

INDUSTRIAL PREPARATION OF NITROGEN

Nitrogen can be produced industrially by the liquefaction of air and then the fractional distillation of the liquefied air

PHYSICAL PROPERTIES OF NITROGEN

Nitrogen is a colorless

Odorless

Tasteless

Pure nitrogen is slightly less dense than air (i.e. its density is almost the same as that of air). It is slightly soluble in water

It has a melting point of about -210

It has a boiling point of about -196

CHEMICAL PROPERTIES OF NITROGEN

Nitrogen is naturally a non-reactive gas due to its covalent triple bond between its atoms. At high temperature and pressure

Nitrogen combines with non-metals

N2+3H2 (reversible) 2NH3

N2+O22NO

The second reaction occurs naturally when lightning flashes at high temperature.

Nitrogen also reacts with metals

3Mg(Red hot)+N2Mg3N2

Mg3N2+6H2O (warm)3Mg(OH)2+2NH3

USES OF NITROGEN

Nitrogen is used industrially to manufacture ammonia

It is also used to manufacture cyanide

It is an essential component in the NPK fertilizer (i.e. Nitrogen, Phosphorus and Potassium)

Liquid nitrogen is used as a cooling agent

Nitrogen is used as a carrier gas in gas chromatography due to its inert nature

It is also used as a preservative as it prevents rancidity in packaged foods

COMPOUNDS OF NITROGEN

AMMONIA

This is a hydride of nitrogen that is produced naturally when nitrogenous compounds decay in the absence of air or by the action of heat or purifying bacteria

LABORATORY PREPARATION OF AMMONIA

Ammonia can be prepared by heating an ammonium salt with a non-volatile base (or strong alkali) like slaked lime [Ca(OH)2]

Nh4Cl+Ca(OH)2CaCl2+2NH3+2H2O

NH4SO4+Ca(OH)2CaSO4+2NH3+2H2O

Ammonia can also be prepared in the laboratory by the hydrolysis of nitrides

Li3N+3H2O3LiOH+NH3

Ca3N2+6H2O3Ca(Oh)2+3NH3

Popular drying agents conc. H2SO4 and fused calcium chloride cannot be used to dry ammonia because they react with the ammonia to produce fertilizer. The most suitable drying agent for ammonia is CaO popularly known as quicklime

INDUSTRIAL PREPARATION OF AMMONIA

Ammonia is obtained industrially by the Haber Process. This involves the direct combination of nitrogen in the ratio 3:1

N2+3H2

For more yield of ammonia,

Air has to be present for nitrogen to be present.

Natural gas is needed to make hydrogen

Steam is needed to make hydrogen and to generate high pressures

A temperature of about 450C is needed

A pressure of about 200atm

The catalyst employed is finely divided ion (mixed with alumina)

PHYSICAL PROPERTIES OF AMMONIA

It is a colorless gas with a chocking pungent smell. It smells like urine

It is less dense than air

It is a poisonous gas and it affects respiratory muscles

This is the only (alkaline) gas than turns red litmus paper to blue

It is very soluble in water forming aqueous ammnia

It has a boiling point of -77.8C

It has a melting point of -34.4C

It does not burn neither does it support combustion

CHEMIACAL PROPERTIES OF AMMONIA

Reaction with air (oxygen): Ammonia burns in air to produce water vapor and nitrogen. In excess air, using platinum as the catalyst, NO is formed

4NH3+3O22N2+6H2O

4NH3+5O24NO+6H2O

Reducing agent: Ammonia is not a strong reducing agent. It reduces (heated) copper (II) oxide to copper while itself is oxidized to water and nitrogen

2NH3+3CuO3Cu+N2+3H2O

NH3 also reduces chlorine to hydrogen chloride gas and nitrogen

2NH3+3Cl26HCl+N2

6HCl+6NH36NH4Cl

Overall reaction

3Cl2+8NH3NH4Cl+N2

If Cl2 is in excess

NH3+3Cl2NCl3+3HCl

NCl3 is an explosive and it is an oily liquid

Reaction with CO2: (at 150C and 150atm, urea is formed)

2NH3+CO2(NH2)2CO+H2O

As a base: Ammonia is a weak base. It reacts with acids to form ammonium salts.

2NH3+H2SO4(NH4)2SO4

NH3+HClNH4Cl

As a precipitating agent: Aqueous ammonia (ammonia in water) precipitates the insoluble hydroxides of metals from solution of their salts

Pb(NO3)2+2NH3+H2OPb(OH)2 (white)+2NH4NO3

CuSO4+2NH3+2H2OCu(OH)2 (blue)+(NH4)2SO4

FeCl3+3NH3+3H2OFe(OH)3 (reddish-brown)+3NH4Cl

ZnSO4+NH3+H2OZn(OH) (white)+(NH4)2SO4

Some metal hydroxides such as copper and zinc will dissolve in excess ammonia solution to form complex ions

Reaction with HCl: Ammonia reacts with HCl to form a dense white fume of ammonium chloride

TEST FOR AMMONIA

Ammonia is the only gas that turns red litmus paper to blue

Ammonia forms a dense white fume with HCl

USES OF AMMONIA

It is used to manufacture fertilizer (nitrogenous fertilizers)

It is used to remove temporary hardness from water

It is used in laundry to remove oil and grease stains

It is used in the production of nylon

Aqueous ammonia was formerly used as a refrigerant but is now replaced by a less toxic substance called fluoro carbon (CF2)

It is used in the manufacture of HNO3 and Na2CO3 by the SOLVAY PROCESS

AMMONIA FOUNTAIN EXPERIMENT

DINITROGEN (I) OXIDE (N2O)

This is called laughing gas. It is a colorless gas with a pleasant smell. Neutral gas an oxide of nitrogen that is neutral to litmus paper.

It is fairly soluble in cold water and it is denser than air

It also supports combustion

PRODUCTION OF N2O

(NH4)2SO4+KNO32NH4NO3+K2SO4

NH4NO3N2O+H2O

TEST FOR N2O

It rekindles a glowing splint and

USES

They are used to cause laughter

Used as inhalational amesthetics

NITROGEN (II) OXIDE

This compound is not known in nature because it is highly unstable

It is slightly denser than air

It supports combustion but however doesn’t rekindle a glowing splint

PREPARATION OF NO

3Cu+8HNO33Cu(NO3)2+4H2O+2NO

NO supports the combustion of burning materials whose flames are not hot enough to decompose them and supply free oxygen

NITROGEN (IV) OXIDE

This is reddish brown gas that has a pungent smell. It is denser than air. It is an acidic gas and is often described as a mixed anhydride because it produces two acids when dissolved in water.

It supports combustion but it does not burn in air

Production of NO2

2Pb(NO3)22PbO+4NO2

CHEMICAL PROPERTIES OF NO2

As a mixed anhydride: NO2 dissolves in water to produce two acids which are HNO3 and HNO2 (because it’s a mixed anhydride)

2NO2+H2OHNO3+HNO2

Reactions with alkalis: NO2 reacts with alkalis to give a mixture of the corresponding salts of the acids

2NO2+2NaOHNaNO3+NaNO2+H2O

When NO2 is cooled, its brown color fades to pale yellow. Here, a dimmer substance which is N2O4 is formed from two molecules of NO2. If heated to about 150C, the dark brown color is restored

Between 20 and 50 degrees

N2O42NO2 (reversible)

Lead Trioxonitrate (V): [Pb(NO3)2]: This is the only nitrate that can be used to produce NO2 in the laboratory because it does not contain water of crystallization.

POTASSIUM NITRATE

This is used mainly as gun powder. Gun powder is a mixture of charcoal, sulfur and KNO3

TRIOXONITRATE (V) ACID

Pure HNO3 is a colorless liquid with a chocking smell

Impure HNO3 turns yellow due to the dissolution of NO2. The impurity can be removed easily by bubbling air through the acid solution

HNO3 is the only acid that does not liberate H2 on reaction with metal (zinc) because it is a powerful oxidizing agent.

The chemistry of HNO3 is highly complicated being a strong acid and a powerful oxidizing agent simultaneously

100% conc HNO3 liberates NO2 when it reacts with copper

4HNO32H2O+4NO2+O2

50% conc HNO3 liberates NO with copper

HNO3 does not rect with aluminium at any concentration. This is why aluminium containers are used to transport HNO3 (in the laboratory)

Pure HNO3 boils at 86C and melts at -47C

It is also highly corrosive

It has a density of 1.5gcm

CHEMICAL PROPERTIES OF HNO3

POWERFUL OXIDIZING AGENT

C+4HNO3CO2+2H2O+4NO2

S+6HNO3H2SO4+2H2O+6NO2

P+5HNO3H3PO4+H2O+5NO2

I+10HNO32H

3Cu+8HNO33Cu(NO3)2+4H2O+2NO

4Mg+10HNO34Mg(NO3)2+3H2O+NH4NO3

CU+4HNO3Cu(NO3)2+2H2O+2NO2

STRONG ACID

NITRATION REACTIONS

INDUSTRIAL PREPARATION OF HNO3

HNO3 is produced on a large scale by the Ostwald process by the catalytic oxidation of ammonia

In the preparation of, only glass apparatus should be used as HNO3 attacks rubber easily.

Nitrate ions (NO3-) i.e. compounds containing nitrates can be tested using the brown ruing test

All group V elements have a general electronic configuration … ns2np3 where n>1

Elements in group V are nitrogen, phosphorus, Arsenic (As), antimony (Sb) and Bismuth (Bi)