**CHEMISTRY**

**MATTER**

Matter is anything that has mass and occupies space. It should be noted that mass is different from weight. Mass is the quantity of matter (or stuff) contained in a body while weight is the force or pull with which the earth exerts a body towards the centre of the earth.

Matter can be divided into three major categories:

Types

Particles

States of Matter

**TYPES OF MATTER**

Elements: These are distinct substances that cannot be split-up into smaller or simpler substances. They can be symbolized by a capital letter, a capital letter and a small letter. The symbols can also be gotten from the Latin and Greek names.

Compounds: A compound that are composed of two or more elements. A compound is formed when these elements are chemically joined. They are usually represented by chemical formulae

Mixtures: These are substances that have an indefinite composition. A mixture is formed when two or more constituents are physically combined

CLASSIFICATION OF ELEMENTS

Based on natural radioactivity, elements can be classified into

Radioactive elements: These are elements that undergo spontaneous decay or degradation followed by the emission of radiations [such as alpha, beta and gamma radiations]. Examples of these elements include Uranium (U), Francium (Fr) and Polonium etc.

Non-radioactive elements: These elements are opposite to the radioactive elements.

Based on the properties, elements can be classified into

Metals: These are elements that ionize by electron loss. Generally, metals ionize as follows

Metals are naturally reducing agents since they undergo oxidization in the ionizing process.

Non-metals: These element ionize by electron gain. They are naturally good oxidizing agents and they undergo reduction

Generally, non-metals ionize as follows

Metalloids: These elements are in between metals and non-metals. They share properties between metals and non-metals

PHYSICAL DIFFERENCES BETWEEN METALS AND NON-METALS

|  |  |
| --- | --- |
| METALS | NON-METALS |
| They are ductile | They are non-ductile |
| They are malleable i.e. they have the property of malleability | They are non-malleable |
| They are usually lustrous (i.e. they have silvery colors) and can be polished | They cannot be polished |
| They are mostly hard solids although sodium is a soft sold and mercury is a liquid metal | They are mainly gases though some are solids and liquids |
| They are sonorous (make sound note) therefore can be used for making musical instruments | They are mostly non-sonorous |
| They are moderately dense though some of them are highly dense (for example iron) | They have low densities though some of them are heavy (diamond) |
| They are good conductors of heat and electricity | They are bad conductors of heat and electricity except graphite. |

CHEMICAL DIFFERENCES BETWEEN METALS AND NON-METALS

|  |  |
| --- | --- |
| METALS | NON-METALS |
| They ionize by electron loss and they undergo oxidation | They ionize by electron gain and undergo reduction  Or |
| They are reducing agents | They are oxidizing agents |
| They are electropositive | They are electronegative |
| They combine with chlorine to form electrovalent chlorides e.g. | They combine with other non-metals to form covalent compounds |
| They from metallic hydrides that are ionic | The acid oxides are acid anhydrides because they dissolve in water to form acidic solutions |
| They replace hydrogen in acids if they are more electropositive than hydrogen | They combine with oxygen to form neutral oxide |
| Most metals combine with oxygen to form basic oxides e.g.  Some oxides are amphoteric (i.e. thay can act as both acids and bases) e.g. etc. | They form mainly acidic oxides e.g. or neutral oxides |

CLASSIFICATION OF COMPOUNDS

Based on the behaviour when exposed to the atmosphere, they can be classified into

Deliquescent compounds: These are compounds that absorb moisture from the atmosphere and form solutions of the compounds

Hygroscopic compounds: These are compounds that absorb moisture from the atmosphere and become wet. The solid hygroscopic compounds form pasty substances and not solution while the liquid ones become diluted.

Efflorescent compounds: These are compounds that give out their water of crystallization to the atmosphere at ordinary temperatures. Examples include Sodium trioxocarbonate (IV) decahydrate

Based on the nature and properties, compounds are also classified as acids, bases and salts

CLASSIFICATION OF MIXTURES

Based on the number of phases coexisting mixtures are classified into

Homogeneous mixtures: These are mixtures that have their constituents existing in just one phase

Heterogeneous mixtures: These mixtures have their constituents existing in two or more different states.

Based on the size of the particles, mixtures are classified into

Solutions: These are homogeneous mixtures of solution and solvents. The solute (substance dissolved to make the solution) could be solid, liquid or gas. The same applies for the solvent (dispersing substance that allows the solute to go into solution)

Suspensions: These are mixtures of small insoluble particles of a solid in a liquid or gas e.g. chalk particles in water

Colloids: These are homogeneous mixtures of two phases (the dispersed phase and the dispersion medium). The dispersed phase has comparison which is analogous to the solute of a true solution while the dispersion medium is analogous to the solvent e.g. starch in boiling water

DIFFERENCES BETWEEN COMPOUNDS AND MIXTURE

|  |  |
| --- | --- |
| COMPOUNDS | MIXTURES |
| They are pure substances | They are impure substances |
| They are chemically combined | They are physically combined |
| They are homogeneous substances | They may be homogeneous or heterogeneous |
| They have chemical formulae | They have no chemical formulae |
| They are groups of elements | They may be group of elements or compounds |
| They are separated chemically | They are separated physically |

SEPARATION TECHNIQUES

Since mixtures are usually impure and are physically combined, they can be separated by physical means. These physical methods are called separation techniques.

Sieving: This is a method of separating solids of different sizes. The instrument used for sieving is called a sieve. The mixture of solid particles is placed on a sieve (with a mesh of a particular size of holes). The particles that are smaller the size of the holes of the mesh will pass through and the ones that are bigger that the size of the holes will stay in the sieve

Decantation: This is a method used to separate a liquid and a denser solid which separate into two layers on standing. For example, a mixture of sand and water

The upper layer (the liquid) is carefully poured into another container (through a decanting tube) leaving the solid residue in the container of the mixture. This method is quick but somewhat inaccurate (or not completely pure)

Filtration: This method is used to separate mixtures of liquids and insoluble solid (particles). A filter paper is used in this process. A funnel is also used and two containers i.e. the container of the mixture and another empty container. The funnel is placed in the empty container and the filter paper is folded into the shape of a cone and placed into the funnel. The mixture is poured into the filter paper. The liquid drains into the empty container while the solid residue is left on the filter paper.

The solid residue is known as residue while the liquid is known as filtrate. This method can be used to separate a mixture of salt in water.

Evaporation: This method is used for separating a dissolved solute from a solution by heating the mixture to dryness so that the liquid portion will evaporate thereby leaving the solid solute e.g. a salt solution. The mixture is poured into an evaporating dish and heated gently over a steam bath or sand bath on the (Bunsen) flame. The liquid is driven off (or evaporated) and the salt is left behind in the dish. The reason a sand bath is used is to reduce the rate of evaporation and prevent the solid from burning.

Distillation: This method is used to recover a solvent from a solution. In other words, it is a process of vaporizing a liquid and then condensing the vapour. The solution is heated in a flask to vaporize the solvent. The vapour passes along the (Liebig’s) condenser is cooled by circulating water in its outer jacket. This method is used to separate two liquids of different boiling points e.g. Ethanol and water. Ethanol boils at 78 while water boils at 100

Fractional distillation: This method is used to separate two or more mixed with liquids with close boiling points into its component fractions. The fractional distillation apparatus is the same as that of distillation except that it has a fractionating column (which is introduced between the distilling flask and the condenser). The fractionating column is packed with glass beads. This is where separation takes place. The upper part is at lower temperature and the lower part at higher temperature. Vapours with high boiling points condense as they get to the upper part and flow back into the distilling flask e.g. Separating Petroleum

Crystallization: This is used to separate salts which easily decompose on heating from their solutions. It’s phenomenon by which crystals are formed from a super saturated solution based on the differences in melting point of the substances. The salt crystals obtained are pure and they contain water of crystallization e.g.

This process takes place where the impurity in the substance is insoluble in the solvent. The impure solid is dissolved in a suitable solvent, leaving the impurities as an insoluble residue. The solution is filtered and heated to drive away some of the solvent until the solution becomes saturated. When the saturated solution cools, crystals of the solute become to form.

Fractional Crystallization: This separation technique is used to separate a mixture of several soluble salts. The technique generally involves a heating and a cooling process and it takes advantage of the solubility of the different salts in the solution. The entire salt mixture or solution is first heated to a very high temperature and this is then followed by a cooling process on attaining the solubility temperature of a particular salt while cooling, the salt crystallizes out of the solution and it is filtered off.

If the purity of a salt or a solute is of utmost importance, the technique of crystallization is applied on extracting the solute.

If the purity of a salt or a solvent is of utmost importance, then the technique of condensation is applied in extracting solvent.

Precipitation: In this method, two liquids and a solid are involved. The solid is soluble in one and insoluble in the obobother. The solid dissolves in the one in which it is soluble and the one in which it is insoluble is added to the mixture. The solid is no longer dissolved and turns back into its solid form and floats on a mixture. The solid residue is called the precipitate. This can also be done by mixing two different solutions which react to produce one soluble compound and an insoluble substance called the precipitate.

Sublimation: This is the direct change of solid to gas or gas to solid. In this method a mixture of solids that sublime and the ones that don’t is used. Examples of solids that sublime are Solid Iodine, Camphor, Solid ammonium chloride, Naphthalene etc.

This separation can be used to separate a mixture of ice (which doesn’t sublime) and camphor (which sublimes).

Magnetic Separation Method: This method is used to separate a magnetic substance and non-magnetic substance. A magnet is passed over the mixture the substances that can be attracted by the magnet are then attracted and the ones that can’t be attracted are left behind

Centrifugation: This involves the use of a device called a centrifuge to separate a mixture of two miscible components with different densities. Centrifugation is used for example in chemical laboratories to separate blood samples into plasma and blood cells. It should be noted that a centrifuge is a machine or device that has the ability to spin test tubes at a very high speed.

Equal volumes of the mixture to be separated are laced in all the test tubes contained in the centrifuge. While spinning, components of the mixture with higher densities (blood cells) are thrown to the bottom of the tube too form Pellets while the components with lower densities (The plasma) get suspended at the top of the tube to form what is called the Supernatant.

Chromatography: This is a method of separating the constituents of a mixture by taking advantage of their different rates of movement in a solvent over an absorbent medium. It is based on the principle that if a fluid containing a number of substances is allowed to pass through an absorbent medium, the different substances in the fluid may travel at different rates and be separated. Solutes which are very soluble in the solvent move up at a faster rate than those which are not so soluble.

There are different types of chromatography which include paper chromatography, gel, column, gas, thin-layer, ion-exchange chromatography etc. The most common and the easiest of these is the paper chromatography.

PAPER CHROMATOGRAPHY

The paper chromatography technique is a separation technique that is used to separate a mixture of several soluble component based on their rate of migration or movement. Mixtures which are to be separated by this method are usually colored mixtures such as paint, dye, chlorophyll etc.

A paper chromatographic setup consists of two phases which are

Stationary phase: This is a strip of chromatographic paper (or alternatively a filter paper) as the case may be

The mobile phase (the moving phase): This is the mixture of ethanol and water in the ratio 3:1

The mixture to be separated (e.g. ink) is spotted close to one end of the chromatographic paper after which the paper is then gently suspended vertically inside the chromatographic glass vessel and then clipped to the lid of the vessel which already contains the mobile phase. The entire setup is allowed to stand for a couple of minutes and during this process, the vapour of the mobile phase is seen as a shadow and on getting the components based on their rate of emigrational movement. The component with the fastest rate of movement emerges first and vice-versa. Solutes which are weakly absorbed by the absorbent medium are easily re-dissolved by the ascending solvent and travel quickly up the absorbent medium (chromatographic paper)

RETENTION FACTOR VALUE (RF-VALUE)

This is a value in the form of a ratio that is used to compare the rate of movement on a chromatographic paper. It is expressed as a ratio of the distance moved by the separated component to the distance moved by the solvent or the mobile phase (solvent front) i.e.

NOTE THE FOLLOWING

A mixture of a solution of and sand can be separated by filtration. Since a solution is a mixture of solute (salt) and solvent (water) and sand is not soluble in water

Chlorophyll and ink can be separated by chromatography

Iodine and Ammonium chloride sublime

A mixture of salt, ammonium chloride and barium sulphate can best be separated by sublimation followed by the addition of water then filtration

Sodium chloride can be obtained from brine by evaporation to dryness

Nitrogen can be gotten from liquid air by fractional distillation

TEST FOR PURITY

The purity of a substance (mixture, compound or element) can be checked based on certain characteristics

Boiling point: The boiling point of a substance is the temperature at which the substance will begin to boil. For example, the boiling point of water is 100 degrees Celsius.

A pure liquid boils at a fixed (definite) temperature while an impure one boils at a wide range of temperature. Normally, the presence of solid impurities in liquids increases the boiling point and therefore the liquid takes more time to boil.

Melting point: This is the temperature at which a solid melts. A pure solid has a fixed melting point and vice-versa. The presence of impurities usually lowers the melting point of a solid

Density: A pure substance (in any state of matter) possesses a fixed density and vice-versa

Paper Chromatography: A pure substance produces only one spot while an impure substance produces multiple spots.

PARTICLES OF MATTER

The particles of matter are the atoms, molecules and ions

An atom is the smallest particle (or part) of a substance (or an element) that can take part in a chemical reaction. Atoms are the smallest possible particles of an element that could exist and still possess the chemical properties of that element

A molecule is the smallest particle of a substance that is capable of independent existence and still retains its chemical properties of that substance. A molecule contains atoms. The atoms are chemically joined. The number of atoms in the molecule of an element/compound is referred to as Atomicity. The number of molecules of an element/compound is represented by a coefficient on the left hand side of the element while the number of atoms is represented by a subscript on the right e.g.

Ions: These can be defined as atoms carrying charge e.g. etc. Ions are formed when atoms gain or lose electrons.

HISTORY OF THE ATOM

The word atom comes from a Greek word which means something that is indivisible

This is the Greek word for cut (tomos). So a-tomos is something that cannot be cut. The Ancient greeks would ponder the nature of things and they considered cutting something in half enough times till it was a single unit that couldn't be cut again. In other words, the greeks theorized about the Atom in 500 BC.

In 400BC, Democritus was an ancient Greek philosopher who lived from 460-360BC. He claimed that there are various basic elements from which all matter is made. He proposed that if we continue to cut a particular matter, a time will come when we eventually end up with the “uncutable” particle. He named that particle the atom. The Greek word “atomos” means not able to be divided i.e. indivisible. He concluded that there was a limit to how far you could divide matter. You’ll eventually end up with a piece of matter that cannot be cut. There are various basic elements from which all matter are made.

What he proposed:

Atoms are small, hard particles

They are made of single materials that are formed into simple shapes and sizes.

They are always moving

They form single materials by joining together

Democritus suggested that all things are made of particles and this concept is still believed up till now.

Aristotle (384 – 322BC) disagreed with Democritus that matter never ends up with indivisible particles. He said it can be divided further. Aristotle’s idea became more popular than that of Democritus. At that time, Democritus’ ideas were rejected by leading philosophers for thousands of years

In 1808, An English scientist (or British chemist and teacher) John Dalton (1776-1884) came up with his own atomic theory. He brought back the idea of Democritus 2000 years after. In late 1700s, scientists learned that elements combine in specific ratio (based on mass) to form compounds. Dalton used actual elements to study elements join together to form new substances. He introduced his idea in 1803

His theory states as follows

All substances are made of atoms which are small particles that cannot be divided, created or destroyed

Atoms of the same element are exactly alike and Atoms of different elements are different

Atoms of elements combine in simple whole number ratios to form chemical compounds (law of definite proportion)

In chemical reactions, atoms are combined, separated or rearranged but never change into atoms of other elements

This explanation also allows us to use chemical equations to describe chemical reactions.

John Dalton’s atomic theory was generally accepted because it explained the conservation of mass, definite proportion, multiple proportion and other observations. Although exceptions to Dalton’s theory are now known, his theories have endured very well with modifications throughout the years.

Dalton was wrong about all elements of the same type being identical. We now know that atoms of the same element can have different neutrons hence, these are called isotopes.

Based on the new discoveries, Dalton’s proposal on solid indestructible atom was abandoned. The discovery of electrically charged particles gave clues that led to the modern theory of atoms

In 1830, Michael Faraday (a British Physicist) made one of the most significant discoveries that led to the idea that atoms had an electrical component. Faraday placed two opposite electrodes in a solution of water containing a dissolved compound. He observed that one of the elements of the dissolved compound accumulated on one electrode and the other was deposited on the

Michael Faraday discovered electromagnetic induction in 1831

In 1833, Michael Faraday (a British Physicist) and Sir Humphry Davy (the inventor of the davy lamp and a very early form of the arc lamp) in their experiments on electrolysis showed that atoms are electrical in nature and may possibly consist of subatomic particles.

In 1879, Sir William Crooks, studied the effects of sending an electric current through a gas in a sealed tube. The tube had electrodes at both ends; and a flow of electrically charged particles moved from one of the electrodes. This electrode was the cathode and the particles were known as cathode rays. These particles were firmly believed to be negatively charged particles

On the first of March,1896, French physicist Henri Becquerel discovered radioactivity. It was an accidental discovery. He just opened his drawer and discovered spontaneous radioactivity.

Marie Currie discovered radon and polonium

In 1897, J.J. Thompson performed experiments on discharge tubes (called the cathode ray experiment) which led to the discovery of the electrons (cathode rays) as a subatomic particle. Based on these experiments the following characteristics of electrons were discovered.

Electrons move in a straight line normal (perpendicular) to the cathode and can cast the shadow of an object placed along their path.

They possess kinetic energy and so can cause the motion of a mechanical wheel placed along their path.

They are negatively charged and therefore attract positive charges and repel negative charges.

They have a constant value ofas its charge to mass ratio no matter the gas used in the tube or the nature of the materials used in the tube. They are fundamental particles of all atoms.

In 1910, an American physicist Robert Millikan (who won the nobel prize in physics in 1923) performed the oil drop experiment and was able to determine the charge on an electron as. This can be used to determine the mass of an electron

In 1911, a New Zealand Physicist, (and consummate experimentalist) Ernest Rutherford discovered the atomic nucleus using a scattering experiment. He performed the gold foil experiment. In this experiment, he bombarded a thin gold foil with alpha particles (generated from a radioactive source). He found out that some of the alpha particles passed through the foil while a few of them (1 out of 8000) were deflected at different angles and some deflected in the same direction backwards. To explain the observation, he suggested an atomic model (the nuclear model) in which an atom has a small positively charged center (nucleus) where nearly all the mass of the atom is concentrated. Surrounding the nucleus is a large space (extranuclear part) containing the electrons. Further experiments showed that the nucleus is made of smaller particles called protons, neutrons and electrons.

Rutherford in 1899 discovered alpha and beta rays. He set forth the laws of radioactive decay and identified alpha particles as helium nuclei.

Ernest rutherford also discovered the protons

In the July of 1913, a Danish Physicist, Niels Henrik David Bohr (a physics Nobel prize winner in 1922) put forward his own model of the atom based on quantum mechanics originally developed by planck. He assumed the Rutherford’s model and suggested that orbits (shells or energy levels) around the nucleus. The orbits/energy levels are designated by the letters K, L, M, N, O ,P and Q

The maximum number of electrons that can be accommodated by an energy level is where n is the value of the principal energy level. As one moves (outwards) from the nucleus the energy level increases

|  |  |  |
| --- | --- | --- |
| Shell | N |  |
| K | 1 | 2 |
| L | 2 | 8 |
| M | 3 | 18 |
| N | 4 | 32 |
| O | 5 | 50 |
| P | 6 |  |
| Q | 7 |  |

Concepts of Bohr’s model of the atom

An electron in an atom exists or revolves in a circular orbit

Energy of an electron is quantized or has a fixed value

An electron emits energy in the form of radiation when it moves from a higher energy state to a lower energy state

Limitations:

This model cannot explain the more complicated spectra lines observed in spectra other than that of hydrogen

Only hydrogen was used for Neil Bohr’s experiments making it only very acceptable for hydrogen

It cannot be used to explain chemical bonding

In 1914, an English Physicist Henry Gwyn Jeffreys Moseley suggested that the number of protons in the nucleus (atomic number) is a fundamental characteristic of an atom while explaining the results of his x-ray experiment on the elements. He published a paper in which he concluded that the atomic number is the number of positive charges in the atomic nucleus. His discovery revealed the true basis of the periodic table and enabled Moseley to predict confidently the existence of four new chemical elements, all of which were found. Rutherford also realized that protons by themselves could not account for the entire mass of the nucleus. He then predicted the neutral particles that could account for the missing mass.

In 1932, An English physicist, Sir James Chadwick, CH, FRS discovered the predicted neutral particle and called it the Neutron. He said they were elementary particles devoid of any electrical charge. James Chadwick won the Nobel Prize in 1935.

ATOMIC NUMBER

This is also called the proton number. The atomic number of an atom is the number of protons present in the nucleus of the atom. It is denoted by the letter “Z”. The atomic number of an element corresponds to its position on the periodic table. In a neutral atom, the number of protons equals the number of electrons. Thus the atomic number of a neutral atom is also the number of electrons.

Frederick Soddy (1877 – 1956) proposed the ideal of isotopes in 1912 which was close to 30 years after Dalton’s idea. Isotopes are atoms of the same elements, having different masses due to varying numbers of neutrons. Soddy won the nobel prize in Chemistry in 1921 for his work on isotopes and radioactive materials. Approximately 50 years after Dalton’s proposal of atoms, evidence began to accumulate which suggested that atoms might not be a sphere that Dalton considered. These evidences came in the form of electrically charged particles

MASS NUMBER

The mass number of an atom/element is the sum of the number of protons and the number of neutrons in the atomic nucleus of the element. It is denoted by the letter A.

An element can therefore be expressed as

RELATIVE ATOMIC MASS

This is the number of times the average mass of one of the element is heavier that one-twelfth of the mass of one atom of carvon-12 i.e. It is the average mass of the atoms of the element on a scale on which one atom of carbon-12 (which is 12 units)

The Relative atomic mass of each element has been verified accurately with the aid of the mass spectrometer introduced by Aston.

RELATIVE MOLECULAR MASS

This is the mass of one molecule of the element involved. It is a multiple of its relative atomic mass.

Here, n is the number of atoms in one molecule of the element

ISOTOPES

Isotopes are atoms with the same atomic number but different mass numbers. These atoms exhibit the property called Isotropy. Isotopes are formed when there is difference in the number of neutrons. They have the same chemical properties because they have the same number of neutrons. They exhibit different physical properties. This is used to determine the proportion/percentage/fraction by which each of the isotopes of an element occurs in nature. This proportion is called its geonormal abundance or (simply) abundance. From this, the Relative Atomic Mass (RAM) of an element (which is the average mass of all isotopes of an element) can be calculated

Here, n is the total number of isotopes involved

Here, m is the mass number of each isotope

Here,is the abundance of each isotope

ELECTRONIC CONFIGURATION

This is like a way of writing the atomic number. As explained earlier, orbitals are found in orbits (K, L, M,..., Q). Orbits are shells, Orbitals are called sub-shells. It should also be noted that orbitals (sub-shells) can be divided into sub-orbitals. So sub-shells are different from sub-orbitals.

ORBITALS

It should be noted that orbits and orbitals are endless. It’s just that for now, we just use K to Q. Similarly, the number of sub-shells is endless but we just use s, p, d, f, g, h and i.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| SN | Symbol | Name of the orbital | Shape of orbital | Maximum number of electrons |
| 1 | S | Sharp | Spherical | 2 |
| 2 | P | Principal | Dumb-bell | 6 |
| 3 | D | Diffused | Double dumbbell (or Rosette) | 10 |
| 4 | F | Fundamental | Complex | 14 |
| 5 | G |  |  | 18 |
| 6 | H |  |  | 22 |
| 7 | I |  |  | 26 |

From the above, it can be seen that the number of electrons in each sub-shell differs by 4.

ELECTRONIC CONFIGURATION IN TERMS OF SHELLS

It should be noted that the atomic number of an element is also the serial number or the position of the element on the periodic table.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| SN | Element | Element symbol | Atomic Number | Electronic configuration |
| 1 | Hydrogen | H | 1 | K L M N  1 |
| 2 | Helium | He | 2 | K L M N  2 |
| 3 | Lithium | Li | 3 | K L M N  2 1 |
| 4 | Beryllium | Be | 4 | K L M N  2 2 |
| 5 | Boron | B | 5 | 2 3 |
| 6 | Carbon | C | 6 | K L M N  2 4 |
| 7 | Nitrogen | N | 7 | K L M N  2 5 |
| 8 | Oxygen | O | 8 | K L M N  2 6 |
| 9 | Fluorine | F | 9 | K L M N  2 7 |
| 10 | Neon | Ne | 10 | K L M N  2 8 |
| 11 | Sodium | Na | 11 | K L M N  2 8 1 |
| 12 | Mag | Mg | 12 | K L M N  2 8 2 |
| 13 |  | Al | 13 | K L M N  2 8 3 |
| 14 |  | Si | 14 | K L M N  2 8 4 |
| 15 |  | P | 15 | K L M N  2 8 5 |
| 16 |  | S | 16 | K L M N  2 8 6 |
| 17 |  | Cl | 17 | K L M N  2 8 7 |
| 18 |  | Ar | 18 | K L M N  2 8 8 |
| 19 |  | K | 19 | K L M N  2 8 8 1 |
| 20 |  | Ca | 20 | K L M N  2 8 8 2 |

Normally, we are supposed to have more than eight electrons in the shells after the L shell but instead, we have a maximum of eight when writing electronic configuration.

Using this KLMN electronic configuration, we can also determine the number of unpaired electrons.

The number of unpaired electrons is given by the last number of the electronic configuration and when the last number is greater than four, then it is that number minus four.

For example, the number of unpaired electrons in the calcium atom; since the last number in the electronic configuration is 2, then the number of unpaired electrons is 2.

Also, the number of unpaired electrons in oxygen, since the last number of the electronic configuration is six (which is greater than 4) we say 6-4. Therefore the number of unpaired electrons is 2.

Normally, when an atom is not in its ionic form (i.e. it is not charged), it’s number of electrons is equal to the number of protons. However, there are a group of elements that don’t have an ionic state. They have an octet electronic configuration i.e. their electronic configuration ends in eight or a duet electronic structure i.e. it has reached a maximum of two electrons and that is if the maximum number of electrons it can hold is two electrons (i.e. in the K shell)

The octet structure or the duet is a stable state and therefore, elements which are not in this state try to attain this state by either gaining or losing gains.

When they gain or lose electrons, they become ionic in state.

For elements whose electronic configurations end in a number 1, 2 or 3, they lose the unpaired electrons and become positively charged.

For example, the electronic configuration of Magnesium ends in 2. Therefore, it loses the two electrons and becomes positively charged with a charge of +2.

Magnesium in its ionic state is written as

Similarly, for Aluminium, it loses its three unpaired electrons and gets a charge of +3.

Aluminium in its ionic state is written as

For elements whose electronic configurations end in 5, 6 or 7, they gain electrons (which are negative) and therefore become negatively charged. The number of electrons that is required to make the element get an octet shape is the charge it will get and is also the number of unpaired electrons. To get the number of electrons, we subtract the last number of its electronic configuration from eight.

For example, the electronic configuration of oxygen ends in 6. The number of unpaired electrons is 2 and therefore, two electrons are added to give it a charge of -2. Therefore, oxygen it its charged state is given as

AUFBAU PRINCIPLE

This principle states that the sublevel shell are filled before those with higher energies. One important thing about this principle is the sublevels do not fill up in numerical order.

PAULI EXCLUSION PRINCIPLE

This principle states that no two electrons in an atom can have the same set of four quantum numbers (n, l, m and s). The principle indicates that no two electrons in the same orbital can spin in the same direction. It limits the number of electrons that can reside in an orbital to two.

HUND’S RULE OF MAXIMUM MULTIPLICITY

This state “Electrons go into degenerate orbitals of sub level (p d f) singly before pairing commences”. This rule is illustrated in orbital notation of Nitrogen and Oxygen.

7N (1s2 2s2 2p3)

8O (1s2 2s2 2p4)

This rule is useful in determining the number of unpaired electrons in an atom.

An element whose atoms/molecules contain unpaired electron(s) is paramagnetic (i.e. it is weakly attracted to things in a magnetic field)

If the unpaired electrons are much (highly paramagnetic) the atom is ferromagnetic (strongly attached).

An element whose atoms/ molecules contain complete paired electrons is diamagnetic (i.e. it is not attracted in magnetic field). If the unpaired

STATES OF MATTER

These are also called the phases of matter

PROPERTIES OF MATTER

Physical properties: These are properties that are observed to without causing any change in the chemical composition of the matter

Chemical Properties: These are properties of matter that describe the chemical changes (or chemical reactions) that matter undergoes.

Extrinsic properties: These are properties that are not characteristics of any particular type of matter e.g. mass, length and temperature

TYPES OF CHANGE

Physical Change

This change is easier to reverse

No new substance is formed

There is no change in mass of the substance

There is no major heat change except latent heat in change of state

CHEMICAL CHANGE

This change is not easily reversed

New substances are usually formed

There is usually a change in t the mass of the substance

There is usually heat change i.e. absorption or evolution of heat

Examples of chemical change

Burning of wood (or anything else)

Dissolution of active metal in cold water

Passing steam over red hot iron or coke

Slaking lime

Rusting of iron in moist air or aerated water

Heating limestone

Passing a steam of hydrogen over heated copper (II) oxide

Adding water to calcium dicarbide

Action of acid on metals, trioxocarbonates (IV) or alkalis (bases)

Charring of sugar

Fermentation of starch to ethyl alcohol

Redox changes in electrochemical/electrolytic cells

Questions

A small quantity of solid ammonium chloride was heated gently in a test tube, the solid gradually disappeared to produce a mixture of two gases. Later a white cloudy deposit observed on the cooler part of the test tube. The ammonium chloride has said to have undergone? Answer: Sublimation.

Which of the following changes is a physical?

A adding iron fillings to aerated water

B adding sodium metal to water

C cooling a solution of iron(II) sulphate

D. Cooling water to obtain ice

E Burning domestic gas

Answer: D

By means of filtration one component can be obtained pure from an aqueous mixture of sodium chloride and

A Potassium nitrate

B sand

C lead nitrate

D sugar (glucose)

E starch

CHEMICAL COMBINATION

In writing chemical formulae, the knowledge of the oxidation number and symbols of elements are needed.

Rules in oxidation number

All group one elements have the oxidation number of +1

All group two elements have the oxidation number of +2

All group three elements have the oxidation number of +3

Oxygen always has a charge of -2 except in peroxides like Na\_2O\_2 H\_2O\_2 where it has an oxidation number of -1

Hydrogen always has an oxidation number of +1 except in metallic hydrides where it has an oxidation number of -1

The sum of oxidation number of elements in a compound should always equal zero

The oxidation number of an element in its uncombined state is 0

The sum of oxidation numbers of the elements of radicals equal to the charge carried by the radical. A radical is a group of atoms joined together chemically but possess a charge

RULES FOR WRITING CHEMICAL FORMULAE

Write the symbols for the elements and radicals

Write the oxidation numbers as superscripts to the right hand side

Rewrite the symbols interchanging the oxidation numbers and write the numbers as subscripts to the right of the symbols as subscripts and omit the signs.

LAWS OF CHEMICAL COMBINATION

Law of conservation of mass (matter): This law states that mass or matter can neither be created nor destroyed during the course of a chemical reaction but can only be converted from one form to another. This law is usually obeyed by a balanced chemical equation in that under normal chemical conditions, the total mass (or masses) of the reactant (or reactants) is equal to the total mass of the products

For a balanced chemical equation to obey the law of conservation of mass, certain conditions must not occur

None of the products of the reaction must be a gas except the reaction is carried out in an enclosed vessel to prevent the escape of the gaseous produce

The reaction must not involve a great amount of heat loss.

LAW OF CONSTANT COMPOSITION

This law states that “all pure samples of a particular chemical compound contains the same kinds of elements combined in the same proportion by mass”.

If three different students for example decide to prepare a sample of Calcium Chloride with each student using an entirely different method, the two elements that must be present in each sample are calcium and chlorine and their ratio of combination must be 1:2

LAW OF MULTIPLE PROPORTION

This law states that “when two elements A and B are combined to form more than one chemical compound, the several masses of A that combine with the fixed mass of B are in simple whole number ratio”

QUESTIONS

In an experiment 5.00g of a metal gave 10.59g of its chloride. In another experiment, 1.5g of the same metal combined with 1.68g of chlorine. Which chemical law is supported by this statement?

|  |  |  |
| --- | --- | --- |
|  | Chloride 1 | Chloride 2 |
| Mass of chloride | 10.59 |  |
| Mass of metal | 5g | 1.5g |
| Mass of chlorine | 10.59 - 5 = 5.59 | 1.68 |
| Ration of metal to chlorine |  |  |

Since both are equal this is the law of constant composition

In two separate experiments, 0.36g and 0.71g of chlorine combine with a metal ‘X’ to give Y and Z respectively. An analysis shows that Y and Z contain 0.2g and 0.4g of X respectively. What chemical law is supported by this statement?

2.85g of an oxide of copper gave 2.52g of copper on reduction and 1.9g of an oxide gave 1.52g of copper on reduction. Which chemical law is illustrated by the data?

An element E forms the following compounds with bromine EBr, EBr\_3 and EBr\_4. Which chemical law is illustrated by this illustration?

In two separate experiments, 0.81g and 0.36g of chlorine combine with a metal M go give A and B respectively. An analysis showed that A and B contain 0.1g and 0.2g of air respectively. What law is illustrated by this data?

An element M forms two oxides 1 and 2. 10g of each oxide contains 1.947g and 6.962 of X respectively. Calculate the mass of x which combines with 1g of oxygen in each oxide and hence show the chemical law supported by this statement.

If the first oxide has the formula XO, what is the formula of the second oxide?

CHEMICAL BONDING

This is an aspect of chemistry that deals with all the forces of interaction existing between atoms, molecules and ions. Basically, atoms are known not to be stable with the exception of the noble gases which are the only group of elements that are known to be neutrally stable. Therefore atoms exhibit the tendency to go into bonding for them to attain stability.

Bonding can be achieved through the following processes:

Electron transfer (loss or gain)

Electron sharing

There are generally two types of interactive forces or chemical bondings which are

Intramolecular (inter atomic) forces of attraction

Intermolecular forces of attraction

INTRAMOLECULAR FORCES

An interatomic force is an attractive force which exists between the particles of a molecule (i.e. between the atoms). There are three types of intramolecular forces namely:

Electrovalent (ionic) bond

Covalent

Metallic bond

ELECTROVALENT OR IONIC BONDING

This chemical bond is formed as a result of the transfer of electrons from one atom (usually a metallic atom) to another atom (usually a non-metallic atom). It should be noted that for this bond to be formed, the difference in the electronegativity values between both atoms must be greater than one

Show the formation of the following compounds

Sodium Chloride

Calcium oxide

Aluminium chloride

Magnesium chloride

Sodium sulphide

Sodium fluoride

Potassium nitride

COVALENT BONDING

A covalent bond is an interatomic bond which is formed as a result of the sharing of electrons between two atoms or a group of atoms. The electrons shared by atoms can either be equally shared or the shared electrons can be contributed by only one of the participating atoms or group of atoms. In the former case (where electrons are shared equally), the covalent bond formed is called ordinary covalent bond while the latter is called a dative (or coordinate) covalent bond. The type of bond in complex ions is the dative bond

The number of electrons that an atom will need to be stable is the number of electrons it will share in a covalent bond

Show the formation of an ordinary covalent bond in the following molecules

Hydrogen (H\_2)

Oxygen (O\_2)

Nitrogen

Water

Ammonia

Carbon dioxide

Hydrogen sulphide

Hydrazine (N\_2H\_4)

Show the formation of a coordinate covalent bond in the following

Hydroxonium ion (H\_3O^+)

Ammonium ion ()

DIFFERENCES BETWEEN IONIC AND COVALENT BONDS

|  |  |
| --- | --- |
| Electrovalent | Covalent |
| In solution, they form ions | In solution, they form molecules |
| They are usually electrolytes (i.e. they can conduct electricity) | They are non electrolytes with the exception of some organic acids |
| They have high melting points and low boiling points | They have low melting points and high boiling points |
| They are soluble in polar solvents such as water | They are insoluble in polar acids (with the exception of organic acids and alkanols) |
| They are non-volatile compounds | They are volatile |
|  |  |
|  |  |

METALLIC BOND

A metallic bond is a bond which id formed between metals and it occurs as a result of the interaction between the nucleus and the element cloud. In other words it can simply be said that the interaction between ion pairs of electrons in the metal nucleus

INTERMOLECULAR FORCES

These are the forces of attraction which exist between molecules. Unlike the intra-atomic forces, the intermolecular forces are generally weak forces and they are of two types:

Hydrogen bond

Vander Waal’s forces

HYDROGEN BOND

This is formed when hydrogen combines with a highly electronegative atom such as oxygen, nitrogen, fluorine and chlorine etc. or with a highly electronegative group (or radical) such as . Hence, hydrogen bond is known to be present in molecules such as water and in acids

VANDER WAAL’S FORCES

There are two types of vander waal’s forces namely:

Dipole – dipole force of attraction

London dispersion forces

DIPOLE FORCE OF ATTRACTION

This is a force of attraction which is known to exist or occur in polar molecules such as water,

A polar substance is one which is known to possess an equal distribution of charges e.g. water and .

From the way I understand polar substances, if we take the formation of water, we will see that the charge of hydrogen is +1 and the charge of OH is -1 we see that we do not need to multiply these charges to balance them out.

Also taking HCl, the charge of Hydrogen is +1 and that of chlorine is -1 so no need to increase the number of any atom in order to balance the charges therefore the charges are evenly distributed to the compound

LONDON DISPERSION FORCES

The dispersion forces on the other hand are forces of attraction which are known to exist in non-polar molecules e.g. gases such as CO\_2, SO\_2, NO\_2 and CH\_4 etc.

Below are the types of chemical bonds in decreasing order of strength

**STOICHIOMETRY**

CALCULATION OF PERCENTAGE COMPOSITION

The percentage composition of an atom in a compound refers to the mass or the molar mass

Calculate the percentage composition of the underlined components in the following compounds

What is the value of x in the molecular formula, if the percentage by mass of Nitrogen is 8.46%

What is the value of x in the molecular for the anion if the central atom (Y) has an oxidation number of +2

The oxidation number of oxygen is different from its value in water in

Calculation of number of particles

Avogadro’s constant,

What is the number of hydrogen ions present in a solution containing (H = 1, S = 32, O = 16)

How many molecules of oxygen are present in 8g of Oxygen

How many atoms of oxygen are present in 8g of Oxygen gas [O=16]

Calculate the number of Hydrogen ions in 4.9g of [H=1N=14, O=16]

What mass of contains the same number of molecules as 0.8g of CH4

How many molecules of phosphorus are present in 496g of the substance (P = 31)

Calculate the number of chlorine atoms present in 7.45g of KCl (K = 39, Cl = 35.5)

How many atoms of Oxygen are present in 8.8g of (C=12, O=16)

What is the ration of the number of molecules of 4g of Hydrogen to that in 32g of Oxygen?

Which of the following is the same as 24g of Mg

1. 1g of H\_2 molecule b. 16g of O\_2 molecules c. 32g of O\_2 molecules d. 35.5g of Cl\_2

An element A has the electronic configuration . The combination of A with a halogen x can give compounds of the formula

EMPIRICAL AND MOLECULAR FORMULA

The empirical formula of a compound is the simplest formula of the compound which indicates the ratio of the different kinds of atoms present in the compound.

The molecular formula of a compound on the other hand is a formula which processes the actual number of each kind of atom.

Glucose for example has the formula . This is regarded as the molecular formula of glucose because it expresses the actual number of carbon, hydrogen and oxygen atoms in the molecule of glucose. However, this formula can be reduced to its simplest form which is and this is known as the empirical formula of glucose since it indicates the ratio of the combination of the different atoms (i.e. 1:2:1).

The following steps below are followed when calculating the empirical formula of a compound

The mass or percentage composition of each atom in the compound is first expressed

The mole ratio of each atom is then calculated. This is done by simply dividing the mass or percentage composition of each atom by its relative atomic mass

Each mole ratio is then divided by the smallest

The empirical formula of the compound is then expressed as mole ratio of each atom in the compound

To calculate the molecular formula of a compound, its empirical formula is equated together with an unknown variable to the molecular mass of the compound

A compound has the following compositions: 1.33g of Potassium, 1.77g of Chromium and 1.9g of Oxygen. Calculate the empirical formula of the compound (K = 39, Cr = 52, O = 16) Ans:

A hydrocarbon is analysed and found to contain 83.3% carbon by mass. What is the empirical formula of the compound (C = 12, H = 1) Ans: C\_5H\_12

What is the empirical formula of an oxide containing 72% Manganese by mass. (Mn = 56, O = 16) Ans:

40g of a hydrate of CaSO\_4 lost 8.368g of water on heating. What is the formula of the hydrate? (Ca = 40, S = 32, O = 16, H = 1). Answer:

A sample of an organic compound containing carbon and hydrogen burns in excess oxygen to yield 8.8g of CO\_2 and 5.4g of H\_2O. What is the empirical formula of the compound? (C = 12, H = 1)

Answer:

On combustion 0.0065g of an organic compound containing carbon, hydrogen and nitrogen only yielded 0.0146g of CO\_2 and 0.0089g of H\_2O. What is the empirical formula of the compound?

Calculate the molecular formula of an organic compound containing 92.31% carbon and 7.69% hydrocarbon by mass if its relative molecular mass is 78g

What is the molecular formula of cryolite if it contains 32.85% Na, 12.85% Al, and 54.30% Fluorine by mass given that the RMM of the compound to be 210 (Na = 23, Al = 27, F = 19)

Upon heating, 1dm^3 of a monoatomic gas (Y) combines with 1.5dm^3 of oxygen to form an oxide. What is the empirical formula of the oxide?

5.0g of the oxide of a metal (M) gave 4.00g of the metal when reduced with hydrogen. What is the empirical formula of the oxide?

The molar ratio of hydrogen to carbon in an organic liquid compound is 2:1. On evaporation at STP, 0.24g of the compound produced 64cm^3 of the vapour. What is the molecular formula of the liquid volume? (C = 12, H = 1, Volume of gas at STP = 22.4dm^3)

An alkanol containing 60% carbon by mass will have a molecular formula of what?

Formula for an alkanol:

Percentage of carbon

Therefore the formula:

CONCENTRATION QUESTIONS

Formulae to note

5 dm^3 of water was added to a solution of having a concentration of 2.0moldm^-3 and 7.5dm^3 volume. Calculate the resulting concentration of the solution. Answer: 1.2

If 1 litre of 2.2 molar sulphuric acid is poured into a bucket containing 10 litres of water and the resulting solution is thoroughly mixed, calculate the resulting concentration of the acid Answer: 0.2

How many moles of CaCO\_3 are there in 2.5g of compound? Answer: 0.025mol

What is the mass of 2.3moles of sodium? [Na = 23] Answer: 52.9

What is the mass of present in of of solution of the compound? Answer 56g

TITRATION QUESTIONS

The concentration of a solution is directly proportional to the number of moles and inversely proportional to the volume of the substance

For acid base titration

= molar conc. (concentration in) of Acid

= molar concentration of Base

= Volume of Acid (End point)

= Volume of Base (in pipette)

= stoichiometric ratio of acid (number of moles of acid in equation)

= stoichiometric ratio of base (number of moles of base in equation)

PRINCIPLE OF DILUTION

The number of moles of acid or base before dilution is equal to the number of moles of the solute after dilution

But,

Volume of water added to the solution is given as,

For neutralization to occur,

Here, the concentration is in

of a solution containing 1.33g or in requires of a solution of for complete Neutralization. What is the concentration of the acid in? Answer 0.05M

2g of a mixture of and NaCl were dissolved in a standard flask. portions of this solution required of HCl for neutralization. What is the percentage by mass of? Answer: 70.255%

5.00g of a mixture of and liberated 1.32 of on strong heating what is the percentage of in the mixture? [C = 12, O = 16, Ca = 40] Answer: 40%

What volume of will exactly neutralize of NaOH solution

MASS VOLUME RELATIONSHIP

This is an aspect of stoichiometry that deals with the relationship between the mass and volume of reactants and/or products. The first step to solving questions in this aspect is to writing balanced equations after which comparisons are then made.

Comparisons are usually made between:

Mole and mole

Mole and volume

Volume and volume

Mass and Mass

Mass and molar mass

Molar mass and molar mass

Questions

How many grams of water would be formed if 0.4g of Hydrogen reacts with enough quantity of oxygen? Answer: 3.6g

Calculate the volume of nitrogen that will be produced at STP from the decomposition of 9.6g of Ammonium dioxonitrate (III) [H= 1, N = 14, O = 16]. Answer:

How much magnesium is required to react completely with of HCl (Mg = 24)

What is the amount in grams of water produced when 5g of hydrogen and 5g of oxygen react together

Calculate the number of molecules of carbon (IV) oxide produced when 10g of is treated with of of

KINETIC THEORY OF MATTER AND GAS LAWS

The average kinetic theory of a gas is directly proportional to its temperature

The molecules of a gas collide with one another and with the wall of their container (and the collisions with the wall constitute the gas pressure)

Collisions of the gases are perfectly elastic

The actual volume occupied by the gas molecules is negligible compared to the total volume of the container

The cohesive forces between the gas molecules are negligible

Vapor density is defined as the mass of a unit molume of the gas compared to an equal volume of hydrogen

At STP (standard pressure and time), T = 273K and P = 760mmHg

WAYS OF COLLECTING GASES

The method of collecting gases depends on the vapour density.

The vapour density of air is 14.4

Vapour density of a gas is

For gases less dense than air i.e. VD < 14.4

Downward displacement (or movement) of air (also known as upward delivery of gas itself)

For gases denser than air i.e. VD > 14.4

We use the upward displacement of air (also known as the downward delivery of the gas)

The third method is collecting the gas over water if the gas is insoluble in water or it does not react with water e.g.

Compression would cause an increase in pressure hence an increase in temperature

GRAHAM’S LAW OF DIFFUSSION

This law states that “the rate of diffusion of a gas is inversely proportional to the square root of its vapour density (or square root of molecular mass”

Note also that

From this

Questions

The densities of two gases A and B are 0.25gdm^-3 and 1.0gdm^-3 respectively. What is the rate of diffusion of A relative to B?

If 50cm^3 of oxygen diffuses through a porous plug in 10s, how long will it take 100cm^3 of methane to diffuse through the same plug under similar conditions? (H = 1, C = 12, O = 16)

Arrange the following gases in increasing order of their rate of diffusion CO, SO\_2, H\_2S, NO\_2, O\_2

A given volume of SO\_2 diffuses in 60s. How long will it take the same volume of methane to diffuse through the same plug under the same conditions?

If SO\_2 and CH\_4 gases are released at the same time at opposite ends of tubes, calculate the ratio of the rate of diffusion of SO\_2 to that of CH\_4

DALTON’S LAW OF PARTIAL PRESSURE

This law states that “if a mixture of gases which do not react chemically are given, then the total pressure exerted by the gas mixture is the sum of the individual partial pressures of each of the gases.

Given gases A, B and C

The total pressure can also be expressed as

Questions:

A gaseous mixture containing 64g of oxygen and 70g of Nitrogen, exerts a total pressure of 1.8ATM. What is the total partial pressure exerted by oxygen in the mixture?

What is the partial pressure of oxygen gas collected over water at standard atmospheric pressure and 25 degrees C. If the saturated vapour pressure of water is 22mmHg at 25C

A gaseous mixture contains 5 moles of helium, 3 moles of Argon and 2 moles of nitrogen. What is the mole fraction of nitrogen in the mixture?

A gaseous mixture of 0.4mole nitrogen, 0.04mole helium and 0.6mole Argon exert a pressure of 6.3ATM. What is the partial pressure of Argon in the mixture

GAY LUSSAC’S LAW OF COMBINING VOLUME

This law states that “when two gases combine, they do so in volumes which bear simple ratio to one another and to the volume of their product if gaseous provided temperature and pressure remain constant”

Questions

Calculate the volume of oxygen required to burn 500cm^3 of methane completely. Answer: 1000

What is the volume of CO\_2 produced in the above question? Answer: 500cm^3

100cm^3 of hydrogen gas was sparked with 100cm^3 of oxygen at 760mmHg. What is the volume of the residual gas at the end of the reaction if the temperature and pressure remain constant? What percentage of this residual gas by volume is steam?

50cm^3 of hydrogen was sparked with of oxygen at 100C and 1atp. What is the volume of residual gas in each of the following cases

At 100C and 1atp Answer: 75cm^3

If passed through an alkaline solution of pyrogalol Answer: 50cm^3

If passed through a tube containing CaCll\_2 Answer: 25cm^3

20cm^3 of CO are mixed and sparked with 200cm^3 of air containing 21% O\_2. What is the total volume of the resulting gas if all volumes are measured at STP? Answer: 210cm^3

40cm^3 solution of hydrogen gas are sparked with 160cm^3 of oxygen at 100C and 1atp. Calculate the volume of oxygen left after the reaction. Answer: 140cm^3

Note that

. It absorbs Oxygen.

. This absorbs both CO\_2 (NaOH) and Oxygen.

AVOGADRO’S LAW

This law states that equal volumes of all gases under the same pressure and temperature contain the same number of particles.

For example, if 10cm^3 of Hydrogen gas contains 6.02X10^23 particles at STP then an equal volume of any other gas will contain the same number of particles at STP

Given that 10cm^3 of oxygen gas contains 2.0X10^4 atoms at 27C and 1atp, calculate the number of atoms that will be contained in:

of ammonia gas at 27C and 1atp Answer:

of gas at 27C and 1atp Answer:

5cm^3 of at 27C and 1atp Answer:

MORE QUESTIONS ON ALL LAWS

Which of the following is correct?

a) The average kinetic energy of a gas is directly proportional to its temperature b) At constant temperature, the volume of a gas increases as the pressure increases c) Pressure of a gas is inversely proportional to its volume d) Temperature of a gas is directly proportional to its volume e) The collisions of molecules with each other are inelastic

30cm^3 of oxygen diffuses through a porous pot in 7 seconds, how long will it take 60cm^3 of chlorine to diffuse through the same pot, if the vapour densities of oxygen and chlorine are 16 and 36 respectively? Answer: 21s

50cm^3 of hydrogen are sparked with 20cm^3 of oxygen at 100C and 1 atmosphere. The total volume of the residual gas is? Answer: 50cm^3

A gas occupied 30dm^3 at STP. What volume would it occupy at 91C and 380mmHg? Answer: 80dm^3

10cm^3 of CO is sparked with 100cm^3 of air containing 21% of O\_2. If all the volumes are measured at STP, the volume of the resulting gases would be? Answer: 105

On heating, under suitable conditions, 1 litre of a monoatomic gas, X, combines with 1.5litres of oxygen to form an oxide, what is the formula of the oxide? Answer: XO\_3

Under high pressure, real gases do not obey gas laws because they slow down as pressure rises and temperature falls

of oxygen was collected over water at 30C and 752mmHg pressure. What is the volume of dry oxygen at STP? Vapour pressure of dry oxygen at 30C = 32mmHg? Answer

Which of the following statement is an exception in the assumptions of kinetic theory of gases/ a. Gases are composed of many elastic particles B) the particles are of negligible size C the particles are in constant random motion D) the particles are negligible mass E. the particles collide with each other? Answer: D

The best way of collecting chlorine is by \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**ACIDS BASES AND SALTS**

Theories on acids and bases

1. Lavoisier’s Theory: Lavoisier defined an acid as a substance which contains oxygen. He defined it as an oxide of a non-metal which when dissolved in water, produces a solution which turns litmus paper from blue to red.

The problem with this theory is that there are many acids without oxygen. E.g.

Lavoisier died in 1816 and he also discovered oxygen

2. Arrhenius’ Theory: Arrhenius defined an acid as a substance that when dissolved in aqueous solution produces hydrogen ion () as the only positive ion.

He defined a base as a substance which when dissolved in aqueous solution produces () as the only negative ion.

His theory is still accepted today.

The above are bases

3. Bronsted – Lowry Theory: J.N. Bronsted and J.M. Lowry simultaneously but independently put forward their theory based on protons.

They defined an acid as a proton (positive charge) donor

They defined a base as a proton acceptor.

Bronsted-Lowry reaction is a reversible reaction in which the products are also capable of donating and accepting a proton. They are generally known as conjugate acids or bases.

In the reactions above is capable of donating and accepting a proton. Hence, is an example of amphoteric specie

4. Lewis Theory: G.M. Lewis’ theory is based on electronic configuration.

He defined an acid as a substance that is capable of accepting an electron pair.

He defined a base as a substance capable of donating an electron pair.

For a substance to act as a Lewis’ base, it must be deficient of electrons. These are known as **Electrophiles (Electron loving species)**.

For a substance to behave as a Lewis’ base, it must possess a lone pair electron. These are generally known as **Nucleophiles (Nucleus loving species)**.

**ACIDS**

An acid is defined as a substance which when dissolved in an aqueous solution (such as water), produces hydrogen ion (H+) or hydroxonium ion (H3O+) as the only positive ions.

**CLASSES OF ACIDS**

Acids can be classified into two based on their sources namely

1. Organic acids: These are acids that can be obtained from plant and animal sources (organic sources). These acids are usually weak acids. The table below shows a few.

|  |  |  |
| --- | --- | --- |
| Acid Name | Acid source (or other name) | Acid Formula |
| Ethanoic acid (formerly called Acetic acid) | Vinegar | CH3COOH |
| Amino acids | Proteins |  |
| Fatty acids | Fats and oils |  |
| Ascorbic acid (Vitamin-C) | Orange |  |
| Citric acid | Lemon and Lime (more pronounced in lemon and lime than in orange) |  |
| Lactic acid | Milk |  |
| Tartaric acid | Grape |  |
| Boric acid | Soda |  |
| Formic acid | Ant |  |
|  |  |  |

2. Inorganic acids: These are also called mineral acids. They are acids that can be obtained from mineral elements or inorganic sources.

**STRENGTH OF AN ACID**

This is defined as the degree of dissociation of acid molecules into ions when put in an aqueous solution. This is called **ionisation**.

There are two types of ionisation namely:

a. Complete ionisation: This is the type of dissociation that is **irreversible**. **Strong acids ionise completely** in aqueous solutions.

The major strong acids are: , and

b. Incomplete (or partial) ionisation: This is the type of dissociation that is **reversible**. Weak acids ionise partially in solution. Examples are H2CO3, H2SO3, H2S and all organic acids etc.

**CONCENTRATION OF AN ACID**

This is defined as the relative proportion of acid to water in a solution.

A concentrated acid has a higher proportion (amount) of acid to water.

A dilute acid has a higher proportion of water to acid.

**NB**: Acids are diluted in order to reduce their concentration.

Acids can be diluted by adding them to water.

In diluting acids, water should not be poured into the flask containing the acid but rather, the acid should be poured into the flask containing the water while the test tube is in a slanting position. This is because acids have affinity for water and they release a large amount of heat when they come in contact with water.

**BASISCITY OF AN ACID**

This is defined as the number of replaceable hydrogen ions in one molecule of the acid.

**GENERAL PHYSICAL CHARACTERISTICS OF ACIDS**

1. Dilute acids have sour tastes
2. Acids (liquid, gases or molten) turn blue litmus paper red
3. Concentrated acids are corrosive. That is they cause severe burns on the skin.

**GENERAL CHEMICAL PROPERTIES (REACTIONS) OF ACIDS**

1. Reaction with metals: Acids react with active metals to liberate hydrogen gas (except HNO\_3)
2. Reaction with Bases or Alkali: This reaction is called a neutralisation reaction. Acids react with bases or alkalis to produce salt and/or water only.
3. Reaction with carbonates: Acids react with carbonated substances to liberate carbon (IV) oxide (CO\_2)

GENERAL USES OF ACIDS

Acids are useful chemicals used to produce other useful chemicals such as fertilizers

Acids are also used in the laboratory as catalysts

BASES

These are oxides and hydroxides of metals. Most bases are insoluble in water but the ones that are soluble form Alkalis. Like an acid, strong alkalis ionize completely in solution while weak alkalis ionize partially in solution

Examples of Bases:

For the alkalis, we have:

KOH (Strong base)

NaOH (Strong base)

NH\_4OH (weak base)

Ca(OH)\_2 (weak base)

For other bases,

Al(OH)\_3

ZnO

CuO

MgO

From the above, we can see that all alkalis are bases but not all bases are alkalis.

An alkali can also be defined as the basic hydroxide which is soluble in water

PHYSICAL PROPERTIES OF BASES

They have bitter taste

They are soapy when touched

They turn red litmus paper blue

Concentrated Alkalis are also corrosive

CHEMICAL PROPERTIES OF BASES

Neutralization: This is the process in which an acid reacts with an appropriate amount of base or alkali to produce salt and water only. It can also be defined as the combination of and to produce .

Ammonolysis: This is the process where by base or alkali react with ammonium salt to liberate ammonia gas

GENERAL USES OF BASES OR ALKALIS

Bases and alkalis are useful chemicals used to produce dyes, perfumes, crayon and glasses etc.

They are used in the purification of Bauxite

SALTS

A salt is a compound formed when the ionisable, from an acid is replaced by a metallic or ammonium ion.

TYPES OF SALTS

Normal salts: These are formed when strong acids react with strong bases. They are formed chemically when the from an acid is completely replaced by a metal.

Normal salts are neutral to litmus paper. Examples of normal salts are:

Acid salt: This is formed when strong acid reacts with weak base. They are produced chemically when the from an acid is partially replaced with a metal. This means dibasic acids form one acid salt, tribasic acids form two acid salts while monobasic acids do not form acidic salts. Acid salts are acidic to litmus paper and they usually contain hydrogen

Examples of acidic salts are

Basic salts: This is formed when strong bases react with weak acids. They can also be produced when there is insufficient supply of acid needed for the complete neutralization. Basic salts usually contain OH^- and they are alkaline to litmus paper.

Examples include:

Double salts; These contain different types of ions in which two are positively charged and the other is negatively charged. Double salts contain water of crystallization and they are very soluble in water. Alum is a typical example of a double salt.

Examples include:

Complex salts: These contain complex ions and they are usually coloured. They are known as the salts of transition metals

PHENOMENA ASSOCIATED WITH SALT

Efflorescence: This is the process whereby some salts when exposed to the atmosphere lose part or all of their water of crystallization to form an anhydrous or a lower hydrate. Such salts are called efflorescent. Examples include

Deliquescence: This is the process whereby some salts when exposed to the atmosphere absorb water (or moisture) and turn into solution. Such salts are called deliquescent. Examples of such compounds include

Hygroscopy: These hygroscopic substances also absorb water when exposed to the atmosphere. If they are solids, they become sticky or moist. Examples include

PREPARATION OF SALTS

The method used in preparing salts depends on the solubility of the salt in water and its stability to hear. A salt is said to be heat resistant if it does not decompose on heating e.g. Na\_2CO\_3, K\_2CO\_3

All sodium, potassium and ammonium salts are soluble in water

All nitrate salts (NO3^-) are also soluble in water

All chloride salts are soluble in water except AgCl and PbCl\_2. However PbCl\_2 is soluble when hot but reappears when cold. CuCl is insoluble in water but CuCl\_2 is soluble, FeCl\_3 is insoluble but FeCl\_2 is soluble

All sulphate (SO\_4^2-) salts are soluble in water except CaSO\_4 and PbSO\_4. However, CaSO\_4 is slightly soluble in water.

Only sodium, potassium and ammonium carbonates (CO\_3^2-) are soluble in water. Other carbonates are insoluble in water

PREPARATION OF SOLUBLE SALTS

Dilute acids on metals

Dilute acid and alkalis

Dilute acid and insoluble base

Dilute Acid and Carbonates

PREPARATION OF INSOLUBLE SALTS

Double decomposition

Combination Reaction

USES OF AMMONIUM SALTS

NH\_4Cl is used as an electrolyte in dry cell.

NH\_4Cl is also used as flux for soldering

(NH\_4)\_2SO\_4 is used as a fertilizer.

(NH\_4)\_2SO\_4 in a more concentrated form is used as a weed killer

(NH\_4)\_2CO\_3 is used as a nose inhaler to prevent dizziness in the laboratory. In a lower concentration, it is used to manufacture crackers

NH\_4NO\_3 is used in making explosive

DECOMPOSITION OF AMMONIUM SALTS

PH-SCALE

This is a scale that is numbered from 1 – 14 used to measure the acidity and/or alkalinity of a substance. A solution with PH less than 7 is acidic while the one with a PH greater than 7 is alkaline. A solution with a PH of 7 Is neutral (i.e. neither acidic nor basic)

Acidity increases with a decrease in PH while alkalinity increases with an increase in PH

In 1909 Sorensen defined the logarithmic equation of PH. He defined Ph as the negative logarithm of the hydrogen ion concentration to base 10.

Adding log to both sides,

Multiplying through by -1

Take note of this too:

Taking as a constant

Calculate the pH of 0.005moldm^-3 H\_2SO\_4. Answer: 2

Calculate the pOH of 2.5X10^-3moldm^-3 of H\_3PO\_4.

Calculate the pH of 2X10^-4moldm^-3 NaOH

Given that the pH of zobo is 3.4, calculate its OH^- concentration

HYDROLYSIS OF SALTS

This can be defined as the chemical reaction between salt and water. It is also known as the reverse reaction of neutralization. The pH of a salt depends on the strength of the acid and base that combine to produce the salt.

The strong acids are

The strong bases are

DRYING AGENTS

These are the substances that have strong affinity for water and/or moisture. They are used for drying gases in the laboratory. They are also used in a desiccator i.e. a device used for drying gases. They may be deliquescent or hygroscopic.

It should be noted that a drying agent should not be used to dry a gas that it can react with e.g. conc. H\_2SO\_4 should never be used to dry NH\_3 because it can react with NH\_3 to produce fertilizers. The table below shows some common drying agents and the gases they can dry

|  |  |
| --- | --- |
| DRYING AGENT | GASES |
| Conc. H\_2SO\_4 | All gases except NH\_3 and H\_2S |
| Fused CaCl\_2 | All gases except NH\_3 |
| P\_2O\_5 | All gases except NH\_3 |
| CaO (Quicklime) | Suitable for NH\_3 only |
| Silica gel (SiO\_2) | All gases |

VOLUMETRIC ANALYSIS

One of the most important aspects of chemistry is to analyse chemical substances qualitatively or quantitatively.

Qualitative analysis involves the determination of the ions present in a given salt.

Quantitative analysis involves the determination of the molar (or concentration) present in a given substance. Quantitative analysis can be volumetric or gravimetric.

Volumetric analysis involves the volume measurement of the substance

Gravimetric analysis involves the direct mass measurement of a substance.

Volumetric analysis is more used because it is more convenient but it is less accurate than the gravimetric analysis.

TITRATION

This is the process of running an acid from the burette into the base in the conical flask until the acid has been completely neutralized by the base. An indicator is usually added to the base so as to detect the colour change. It should be noted that titration is an aspect of volumetric analysis

INDICATORS

These are weak organic acids or bases that change colour in solution according to the hydrogen concentration (pH)

|  |  |  |
| --- | --- | --- |
| Indicator | Colour in acidic solution | Colour in basic solution |
| Litmus paper | Red | Blue |
| Methyl Orange | Red/Pink | Yellow |
| Phenolphthalein | Colourless | Red |

The choice of an indicator depends on the strength of the acids and bases that combine

|  |  |  |
| --- | --- | --- |
| Strength of acid | Strength of base | Suitable indicator |
| Strong Acid | Strong base | Any indicator |
| Strong Acid | Weak Base | Methyl Orange |
| Weak Acid | Strong Base | Phenolphthalein |
| Weak Acid | Weak Base | No suitable indicator |

The concentration of a solution is directly proportional to the number of moles and inversely proportional to the volume of the substance

For acid base titration

C\_a = molar conc. (concentration in moldm^-3) of Acid

C\_b = molar concentration of Base

V\_a = Volume of Acid (End point)

V\_b = Volume of Base (in pipette)

n\_a = stoichiometric ratio of acid (number of moles of acid in equation)

n\_b = stoichiometric ratio of base (number of moles of base in equation)

PRINCIPLE OF DILUTION

The number of moles of acid or base before dilution is equal to the number of moles of the solute after dilution

But,

C\_i = initial concentration at V\_i

C\_f = final concentration at V\_f

Volume of water added to the solution is given as,

SOLUBILITY

The solubility of a solute in a given solvent is the amount of solute in gdm^-3 or moldm^-3 that will dissolve in 1000cm^3 or 1dm^3 of the solvent to form a saturated solution

Volume in cm^3

Volume in dm^3

SOLUBILITY PRODUCT

We represent each ion or radical by the same letter

BUFFERED SOLUTION

A buffered solution is one that does not bring about a change in pH on the addition or dilution of small amount of alkali.

Buffered solution principle is widely employed in the manufacture of drugs. Buffered solution is prepared by mixing a weak acid and its salt or a weak base and its salt e.g.

**ELECTROCHEMISTRY**

This part of chemistry deals with how electricity can be gotten from chemicals or can be converted to chemical energy. It can also be defined as the relationship between electricity and chemistry.

**TERMS USED IN ELECTROCHEMISTRY**

**1. Electrolyte**: This is a substance that will conduct electricity either in aqueous or molten state and then decompose during the process. Examples of electrolytes include aqueous or molten solutions of acids, bases or salts.

There are two types of electrolytes which are:   
i. Strong electrolytes: These are electrolytes which can **ionize completely** when in aqueous form or in molten form.

E.g. Mineral acids (HCl, HNO3, H2SO4)

Caustic alkalis (NaOH, KOH) and

Salts (NaCl, K2SO4, NH4­Cl etc.)

The conductivity of a strong electrolyte decreases slightly with an increase in concentration.

ii. Weak Electrolytes: These are electrolytes that **ionize partially** in aqueous form or molten form.

E.g. Weak (Organic) acids (HCOOH, CH3COOH)

Weak bases (NH3) and water.

The conductivity of a weak electrolyte decreases greatly with an increase in concentration.

**NB**:

i. Generally, the conductivity of an electrolyte reduces with an increase in concentration.

Where Λ= conductivity

Where c = concentration

ii. The difference between complete ionization and incomplete ionization is that

a. **Complete ionization is irreversible** while **incomplete ionization is reversible**.

**2. Electrodes**: These are materials in the form of metals, wires or plates which help in conveying electrons or current in or out of an electrolyte. There are two types of electrodes namely:

a. Anode: **In electrolysis (an electrolytic cell), this is the positive electrode** but **in an electrochemical cell, this is the negative electrode**. At the **anode**, the process of **oxidation (electron loss)** takes place (in electrolysis).

The anode carries current into the electrolyte and carries electrons out of the electrolyte

b. Cathode: **In electrolysis (an electrolytic cell), this is the negative electrode** but **in an electrochemical cell, this is the positive electrode**. At the cathode, the process of reduction (electron gain) takes place (in electrolysis)

The cathode **carries current away from the electrolyte** and **carries electrons into the electrolyte**.

**ELECTROLYSIS**

This is a part of electrochemistry.

This can be defined as the decomposition of an electrolyte due to the passage of electric current through.

The power source of an electrolytic setup is a battery or a direct current source.

Electrolysis involves three stages namely:

1. Ionisation of an electrolyte and water (if in aqueous form) or the electrolyte alone (if in molten form)

E.g. Aqueous CuSO4 (CuSO4 and water will ionise since it is in aqueous form)

CuSO4 → Cu2+ (aq) + SO42-(aq) [Complete ionisation]

H2O H+ + OH- [Incomplete ionisation]

2. Migration of ions to the electrodes. When an electric current is passed an electrolyte, the electrolyte decomposes into positively and negatively charged ions. If the electrolyte is in aqueous state, water will also decompose into hydrogen ion (H+) and hydroxide ion (OH-).

The **positively charged ions migrate towards the cathode** which is the negative electrode (**Unlike charges attract**). The **positively charged ions** in electrolysis are therefore called **Cations (Cathode ions)**.

The **negatively charged ions migrate towards the anode** which is the positive electrode. These **negatively charged ions** are called **Anions (Anode ions)**.

3. Discharge of ions at the electrode: This discharge involves reduction and oxidation. When ions get to their respective electrodes, they struggle for preferential discharge.

**THE ELECTROCHEMICAL SERIES**

**CATIONS (POSITIVE IONS)**

K+ (Potassium ion)

Na+ (Sodium ion)

Ca2+ (Calcium ion)

Mg2+ (Magnesium ion)

Al3+ (Aluminium ion)

Zn2+ (Zinc ion)

Fe2+

Sn2+

Pb2+

H+

Cu2+

Hg+

Ag+

Au3+

**ANIONS (NEGATIVE IONS)**

OH-

I-

Br-

Cl -

NO3-

SO42-

F-

**FACTORS AFFECTING PREFERENTIAL DISCHARGE**

There are three major factors which affect preferential discharge namely:

1. Position of the ion in the electrochemical series.

For **cations**, **those which are down the series are usually discharged at the cathode in preference to those that are higher the series**. This is because cations which are lower the series are less electropositive (or more electronegative) than those which are higher the series and possess a greater tendency to gain electrons (reduction) since reduction occurs at the cathode.

For **anions**, **those which are higher the series are discharged at the anode in preference to those that are lower the series**. This is because anions which are higher the series are less electronegative (or more electropositive) than those which are lower the series and possess a greater tendency to lose s at the electrons (oxidation) since oxidation occurs at the anode.

2. Concentration: Generally ions which are more concentrated are discharged at the electrodes in preference to those which are less concentrated. However for concentration to be considered as a factor of ionic discharge during electrolysis, the number of ions between the two competing ions **must not exceed three**.

For example, in the electrolysis of a brine (a concentrated form of (NaCl), Cl- is deposited at the anode in preference to OH- because Cl- is more concentrated than OH- and the number of ions in the E.C.S is two which is less than three.

At the cathode, H+ is deposited in preference to Na+ even though is more concentrated. This is because the number of ions between them is more than three in the E.C.S.

3. Nature of the electrode used in the electrolysis process: Generally, electrodes which are used during electrolysis can either be active or inactive (inert).

Inert electrodes do not take part in the electrolytic processes. In other words, they do not determine the discharge of ions during electrolysis examples include carbon and platinum electrodes.

The active electrodes are those which to a large extent determine ionic discharge i.e. they participate in electrolytic processes. An example is the mercury electrode. For example, in an electrolytic process involving an electrolyte of sodium using mercury as the cathode, sodium will always be deposited on the cathode irrespective of every other factor. This is as a result of the affinity of mercury for sodium. Sodium therefore reacts with mercury to form a complex called sodium malgam.

**NB**:

1. Any metal above hydrogen on the activity series is an active electrode.

2. Any time any metal from the activity series is used as the anode, there will be no anodic reaction because the anode will melt into the electrolyte.

3. The charge of oxygen in electrolysis only is always -4.

4. Whenever OH- is deposited on the anode, the reaction is given as

4OH- →2H2O + O2 + 4e-

**FARADAY’S LAWS OF ELECTROLYSIS**

First Law: This law states that

The mass of a substance deposited at an electrode during electrolysis is directly proportional to the quantity of charge flowing through the electrolyte.

Q is the quantity of charge measured in coulombs (C)

But

Where I = Current [measured in amperes (A)]

Where t = Time [measured in seconds (s)]

But

Where M = Molar mass of the element

Where q = the charge of the element

Where F = Faradays constant [measured in farads (F)]

The letter **‘z’** is a constant called the **electrochemical equivalent (E.C.E)**

It can be defined as the mass of a substance deposited when one coulomb of electricity is passed through the electrolyte.

**NB**: One Farad = 96500 coulombs

1F = 96500C

Second Law: This state that

If the same quantity of electricity is passed through two or more different electrolytes, the relative number of moles of the element deposited on the electrode is inversely proportional to the ionic charge of the element deposited.

The letter **‘k’** is a constant.

For two cases,

But

Where m = given mass

Where M = Molar mass

**ELECTROCHEMICAL CELLS**

These are devices used for changing chemical energy into electrical energy.

There are two types of cells namely:

1. Primary cells: These are cells whose reactions (through which electric current is generated) are **irreversible**. They cannot be recharged after they have been used. They have **high resistances** and hence **generate low currents**. They are affected by local action and polarisation.

2. Secondary cells: These are also called **accumulators**. They are the opposite of primary cells.

An electrochemical cell consists of two electrodes of which each has its own electrolyte. The two electrodes are connected by a **salt bridge** or a **porous partition**.

One electrode with its electrolyte is called **a half-cell**. The electrode is an element (mostly metals although it may me non-metals) and the electrolyte is a salt of the element. The element is said to be in contact with its own ions.

A half-cell can be represented as

E(s)/Eq+ (aq) or

An electrode has a quantitative property called the **electrode potential**.

Electrode potential is a measure in volts of the tendency of an atom of an element of the electrode to undergo oxidation and reduction.

M(s) → Mn+ + ne-: Eᶿox

Mn+ + ne- → M: Eᶿred

Eᶿox and Eᶿred of an electrode have the same value but different signs (one is negative while the other is positive).

When all the concentrations of aqueous solutions are 1moldm-3, all gases taking part in the reactions are at 1atm and the temperature is 25⁰C (Room temp), the electrode potential is called the **standard electrode potential**.

The standard hydrogen electrode was assigned a standard electrode potential of 0.000Volts

EᶿH2=0.0000V

The standard electrode potentials of other elements can be determined using the standard hydrogen electrode as a reference.

The standard electrode potential of an element is the potential difference set up between a standard hydrogen electrode and a half-cell.

It is called the electromotive force (emf) when a standard hydrogen electrode is coupled to the electrode of the element.

**NB**:

1. A/An+ is represented as

A→An+ + ne- (oxidation [you don’t need to put this])

2. An+/A is represented as

An+ +ne- → A (reduction)

3. A/An+ // Bn+/B is represented as

A→An+ + ne- EᶿA(ox)

Bn+ +ne- → B

Eᶿcell = EᶿA + EᶿB

E.g. Zn/Zn2+ is represented as

Zn→Zn2+ + 2e-

When calculating the emf of a cell,

i. Write the anodic reaction (Oxidation)

ii. Write the cathode reaction (Reduction)

iii. Make electrons gain equal to electrons lost.

iv. Add the two electrode potentials i.e.

Eᶿcell = EᶿA + EᶿB

**If the emf** of a cell (after calculation) **is positive**, then the **reaction is spontaneous** and the **Gibb’s free energy is negative**.

**If the emf is negative**, **the reaction will not take place**.

The table below shows the standard potentials of elements.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Element | Half-cell reaction (Reduction) | Half-cell reaction (Oxidation) | Standard reduction potential | Standard oxidation potential | Reducing ability | Oxidising ability |
| K |  |  | -2.93 |  | Highest reducing ability (but **can be easily oxidised**) | Lowest oxidising ability |
| Na |  |  | -2.87 |  |  |  |
| Ca |  |  | -2.71 |  |  |  |
| Mg |  |  | -2.37 |  |  |  |
| Al |  |  | -1.67 |  |  |  |
| Zn |  |  | -0.76 |  |  |  |
| Fe |  |  | -0.44 |  |  |  |
| Sn |  |  | -0.14 |  |  |  |
| Pb |  |  | -0.13 |  |  |  |
| H |  |  | 0.00 |  |  |  |
| Cu |  |  | 0.34 |  |  |  |
| Hg |  |  | 0.79  0.85 |  |  |  |
| Ag |  |  | 0.80 |  |  |  |
| Au |  |  | 1.5 |  |  |  |
| OH |  |  | Not an element |  |  |  |
| I |  |  | 0.54 |  |  |  |
| Br |  |  | 1.09 |  |  |  |
| Cl |  |  | 1.36 |  |  |  |
| NO |  |  | Not an element |  |  |  |
| SO |  |  | Not an element |  |  |  |
| F |  |  | 2.87 |  | Lowest reducing ability | Highest oxidising ability |
|  |  |  |  |  |  |  |

**DEFECTS OF PRIMARY CELLS**

1. Local action: This is caused by the presence of impurities on the zinc plate electrode. The impurities on the electrode prevent the free flow of current. Local action can be prevented by **Amalgamation**. Amalgamation is the process of coating the zinc electrode with mercury. Local action can also be prevented by using a pure zinc electrode (which can be gotten from electrolytic process).

2. Polarisation: This is defined as the process whereby hydrogen bubbles produced at the zinc electrode prevent free flow of current. The hydrogen bubbles also increase the resistance of the electrode. The bubbles produce what is known as **Hydrogen Electrode**. These bubbles also create a back emf.

Polarisation can be prevented using chemical substances called **Depolarisers**.

**DIFFERENCE BETWEEN THE ELECTROLYTIC CELL AND THE ELECTROCHEMICAL CELL**

|  |  |
| --- | --- |
| Electrolytic cell | Electrochemical cell |
| Converts electrical energy to chemical energy | Converts chemical energy to electrical energy |
| Electrons are pushed by an outside force such as a battery | Electrons are generated by oxidation at the electrode (anode) |
| Cathode is negative electrode | Cathode is the positive electrode |
| Anode is the positive electrode | Anode is the negative electrode |
| Electrodes are in the same compartment | Electrodes are in separate compartments |
| Salt bridge is not needed | Salt bridge is needed |
| Only one electrolyte is used | Two electrolytes are used |
|  |  |
|  |  |
|  |  |

**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)

**CHEMICAL KINETICS**

This is the branch of physical chemistry that studies the rate of chemical reactions. The rate of a chemical reaction is defined as the speed at which a reaction proceeds. Different reactions occur at different rates. Some reactions are very fast that may occur within micro seconds. Other reactions may also be very slow that they may occur for months or years. Examples of fast reactions include explosions, combustion, neutralization and collision etc. Examples of slow reactions include decaying, fermentation, rusting, oxidation of diamond and organic reactions.

The rate of a reaction is also defined as the rate of disappearance of the reactant or the rate of appearance of the product. In terms of concentration (volume), the rate of a reaction is defined as the decrease in the concentration of the reactant or an increase in concentration (amount) of the product.

**COLLISION THEORY**

This theory assumes that

There must be collision between the reactant particles for a chemical reaction to occur. There are many of such collisions but a small fraction of them lead to the chemical reaction. The small fraction of collision is called EFFECTIVE COLLISION

The minimum amount of energy in which the reactant particles must possess for a chemical reaction to occur is known as the ACTIVATION ENERGY. This means for a chemical reaction to take place, the energy of the colliding particles must be greater than the activation energy

FACTORS THAT AFFECT THE RATE OF CHEMICAL REACTIONS

These are the factors that determine the speed at which a reaction proceeds

Concentration

Temperature

Catalyst

Surface Area Exposed

Light

Nature of the reactants

Pressure for gaseous reactions

CONCENTRATION OF THE REACTANTS

An increase in the concentration of the reactant increases the frequency of collision which leads to a rapid increase in the reaction rate. This means that reactions occur faster at higher concentration that at a lower concentration. For example, conc. H2SO4 liberates hydrogen faster than dilute H2SO4

The SI unit of a rate of a reaction is which is actually the.

RATE LAW

This law states that

The rate of a chemical reaction is directly proportional to the concentration of the reactants raised to some powers

From rate law

Here, x is the order of the reaction with respect to A

And y is the order of the order of the reaction with respect to B

Also, is the overall order of the reaction and this determines the rate of the reaction.

Also, is the molecularity of the reaction. Molecularity of a reaction is defined as the addition of the coefficient of the reactions

The molecularity of a reaction is always a whole number and it is determined stoichiometrically

The order of a reaction can either be zero, fraction, percentage or whole number and it is determined experimentally

From

Here, k in the equation above is the specific rate constant. Its SI unit depends on the overall order of the reaction

Find the SI unit of the specific rate constant of a zero order reaction

The SI unit of k for a zero order reaction is

The SI unit of k for a first order reaction is

The SI unit of k for a second order reaction is

The SI unit of k for a half order reaction is:

The SI unit of k for a 0.125 order reaction is mol7/8dm-21/8s-1

The SI unit of k for a 75% order reaction is mol1/4dm-3/4s-1

RATE CURVE

This is the curve of concentration against time. The rate curve actually depends on the order of the reaction

For zero order, parallel to the x axis

For first order, direct graph

For second order , direct then curve to the left.

RADIO ACTIVE REACTIONS

It should be noted that all radioactive decay follow the first order reactions

From calculus notation

On integrating both sides

At half life

K in the equation above is called the decay constant

TEMPERATURE

Most reversible reactions at ordinary temperatures take months or years to attain equilibrium. However, at higher temperature, equilibrium can be achieved within micro seconds. An increase in temperature increases the frequency of collision and change the content of the reaction which leads to a rapid increase in the reaction rate. This means that reactions occur faster at higher temperatures. Generally, for every degree rise is temperature, the reaction rate doubles. Temperature is defined as the measure of the average kinetic energy of a substance

Arrhenius studied the effect of temperature on the reaction rate

From Arrhenius,

K is the specific rate constant

A is the Arrhenius constant

E is the activation energy

R is the molar gas constant

T is the absolute temperature

From

Adding In to both sides,

In a graph of Ink against 1/t, the y intercept is InA while the slope is –E/R

CATALYST

A catalyst is a substance that alters the rate of a reaction but remains chemically and quantitatively unchanged at the end of the reaction. A reaction that involves a catalyst is called a catalysis or a catalytic reaction. Substances that improve the efficiency of a catalyst are called promoters while substances that decrease the efficiency of a catalyst are called Inhibitors or enzyme poison.

The basic function of a catalyst in a reaction is to change (or increase) the rate of a reaction by lowering the activation energy

GENERAL CHARACTERISTICS OF CATALYSTS

They alter the rate of reactions

They remain unchanged at the end of the reaction

They are specific in action

They also do not affect the product formed

They have no effect on the position of equilibrium

The efficiency of a solid catalyst can be improved by the increase of its surface area.

SURFACE AREA EXPOSED

An increase in the surface area of the reactant increases the frequency of collision which leads to a rapid increase in the reaction rate. For example, powdered marble (which has a large surface area because its particles are freer and can easily interact) liberates CO2 faster than when solid marble when it combines with HCl because of the increase in its surface area. Similarly, sore dust particles of a particular wood burn faster than a log of wood of the same mass.

LIGHT

Some reactions are influenced by the presence of light and they are known as PHOTOCATALYTIC reactions. Examples include

The reaction between hydrogen and chlorine is spontaneous under a bright sunlight but slower under a diffused light

Photosynthesis

Decomposition of hydrogen peroxide

Substitution reaction between methane and chlorine

PHYSICAL NATURE OF THE REACTANTS

The physical nature of the reactants also plays a significant role in determining the speed of a reaction. For example, when gold is dropped into a solution of HCl, there is no evidence of any reaction

When Iron is dropped into HCl, there is a slow evolution of Hydrogen gas

When zinc is dropped into a solution of HCl, there is a fast evidence of a reaction and hydrogen gas is evolved quickly

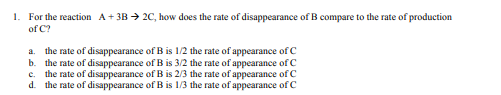
PRESSURE

Pressure affects only gaseous reactions. Generally, an increase in pressure increases the frequency of collision which leads to the rapid increase in the reaction rate

For forward reaction,

For a backward reaction, the products are now the reactants

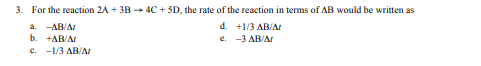
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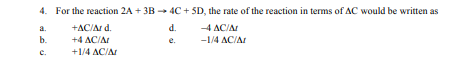
Answer: B



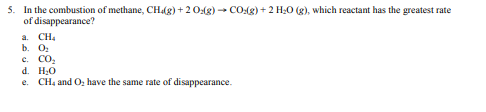
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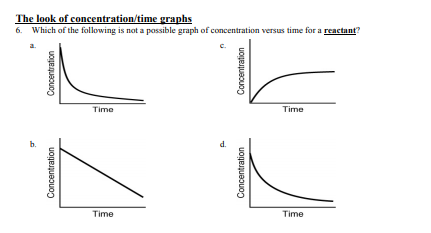
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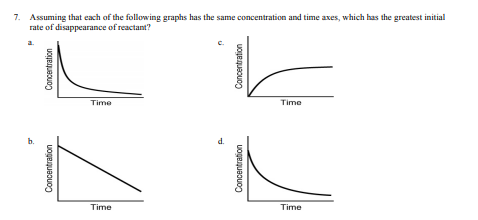
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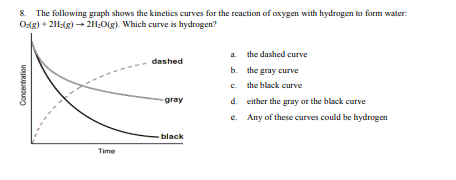
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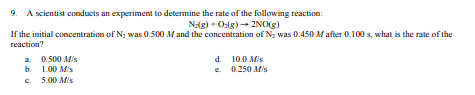
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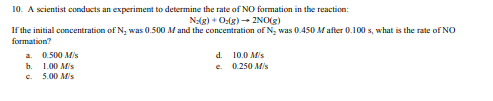
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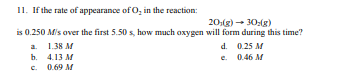
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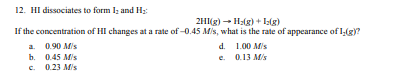
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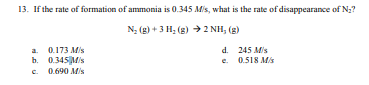
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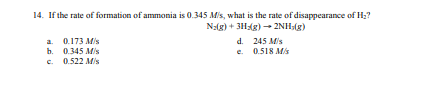
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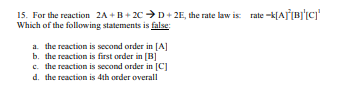
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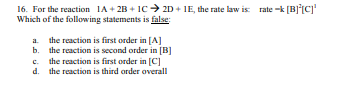
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**THERMOCHEMISTRY**

Why do you have to strike a match to start a fire? Why do some chemical reactions give off heat while other reactions absorb heat? Thermodynamics is the study of energy as it changes from one form to another.

Thermochemistry is the study of energy change in the form of heat associated with chemical reactions. It is an aspect of thermodynamics that focuses on heat evolved or absorbed during a chemical reaction

This is the branch of physical chemistry that studies the heat changes in a reaction. The heat change (enthalpy change) represented as delta-H is usually written alongside the equation. For example,

Heat change is measured in kilo joules per mole

Note: delta H is a state function (i.e. it depends on the initial and final conditions of the reaction). For example, looking at this

If we divide the equation through by 2 (or multiply by half), the delta H will also be divided by 2

And if the reaction is reversed, the heat change also changes

**TERMINOLOGIES USED IN THERMOCHEMISTRY**

System: This is defined as the portion of the universe under study where chemical reactions take place. It is also called a thermodynamic system

Surrounding: This is defined as everything in the universe apart from the thermodynamic system

Boundary: This is defined as the portion or the layer that separates the system from the surrounding

Open System: This is the type of system in which matter and the energy are transferred to the surrounding

Closed System: This is a type of system in which only energy is transferred while matter remains constant

Isolated System: This is a type of system in which both matter and energy remain constant (i.e. they are not transferred to the surrounding

Exothermic (Reaction): This is the reaction in which heat is liberated to the surrounding. It is also defined as a reaction in which the hydration energy is greater than the lattice energy. is negative

Endothermic reaction: This is the reaction in which heat is absorbed from the surrounding into the system

Adiabatic Reaction: This is the reaction in which heat is neither absorbed nor evolved I.e. there is neither heat loss nor heat gain. is equal to zero

Isobaric Reaction: This is a reaction that occurs at constant pressure

Isothermal: Occurs at constant temperature

Isomeric: This is a reaction that occurs at a constant volume. It is also known as an isoconic or an isovolumetric reaction.

Extensive Properties: These are properties that depend on the quantity of matter. The total can be obtained by adding individual constituents together. For example mass, weight, volume, time, heat capacity, energy, work etc.

Intensive Properties: These are properties that are independent on the quantity of matter. The total cannot be obtained by adding the individual constituents. For example temperature, pressure, density, specific heat capacity, refractive index etc.

Hess’ Law OF Constant Heat Summation: This law states that the heat content of a particular reaction is independent on the path taken by the reaction provided that the condition of the initial stage of the reaction is the same as that of the final stage.

For example,

Enthalpy: This is known as the heat content of a reaction. The standard heat of a reaction is defined as the heat energy absorbed or evolved when the chemical reaction occurs between the molar quantity of a substance as represented under standard conditions

Here, m is the stoichiometric ratio (number of moles or atoms) of the product

Also, n is the stoichiometric ratio of the reactant

is the heat change of the reaction

andare the heat changes of the product and reactant respectively

When

STANDARD HEAT OF COMBUSTION

This is defined as the heat energy evolved when one mole of a substance is completely burned in air. The instrument used to mead=sure the heat of combustion is called the bomb calorimeter. For example,

Since in the equation above it’s not one mole of the of C6H6 that is burned, that can’t be the equation of the standard heat of combustion of C6H6

STANDARD HEAT OF FORMATION

This is defined as the heat energy evolved (or absorbed) when one mole of a substance is formed from its constituent elements. This means that for standard heat of formation to occur, the following conditions must be satisfied

One mole of the compound must be formed

The constituents of the product must be the reactants

Only the compound must be the product

The standard heat of formation of all elements in their standard (un-combined) state is zero

Since one mole of H2O wasn’t formed, that can’t be the standard heat of formation of H2O. When modified,

From the reaction above, it can also be deduced that the standard heat of formation of water is also the standard heat of combustion

STANDARD OF NEUTRALIZATION

This is defined as the heat energy evolved (or absorbed) when one of H+ from an acid combines with one mole of OH- from a base to produce one mole of H2O. The standard heat of neutralization of a strong acid and a strong base is constant provided that one mole of H2O is produced

Most standard heats (heat of combustion, formation and neutralization) evolve heat (most especially that of combustion). Therefore

Here, m is the mass, c is the specific heat capacity and Theta is the change in temperature

THERMODYNAMICS

This is the branch of thermochemistry that studies the relationship between heat energy and other forms of energy. In this system, heat energy is highly pronounced and the other forms of energy are regarded as work

FIRST LAW OF THERMODYNAMICS

This law states that

Energy can neither be created nor destroyed but can be converted from one form to another.

This law is also called the law of conservation of energy

is the change in internal energy

Q is the heat energy

W is the work done

When q is positive, the system absorbs heat

When it is negative, the system evolves heat

When w is positive, there is work done on the system

When w is negative, there is work done by the system

ENTROPY

This is defined as the degree of disorderliness or randomness of a substance. Entropy depends majorly on the temperature of the system. Generally, an increase in temperature increases the entropy of a system and a decrease in temperature decreases the entropy of the system. Entropy change is represented as delta S.

Also, as a substance changes from solid to liquid and from liquid to gas, there is an increase in the entropy of the system and when a substance changes from gas to liquid or liquid to solid, there is a decrease in the entropy of the system.

Generally, the change in entropy is given as

Here, T is the absolute temperature

The SI unit of delta S is

The highest degree of entropy can be gotten from a mixture of two (or more) gases

The higher the electronegativity of a substance, the higher the entropy value; for example, amongst the hydrogen halides, hydrogen fluoride has the highest entropy value. Also, fluorine, being the most electronegative element will have the highest entropy while potassium will have the lowest entropy

SPONTANEOUS PROCESS

A reaction is said to be spontaneous if it occurs feasibly (i.e. it takes place without an external influence (mechanical or controlled by man’s influence). For example, the evaporation of a liquid to gas; using a towel to dry the body after a bath; It should be noted that spontaneous reactions are irreversible

SECOND LAW OF THERMODYNAMICS

A reaction is an increase in the entropy of the system and its surrounding.

According to the second law,

Energy Change

Delta S

GIBBS FREE ENERGY CHANGE ()

This is defined as the energy that brings about a chemical reaction. It is also defined as the driving force in a system. It is represented as

The following should be noted

When delta H is positive, t delta S is less than delta H. The reaction is non-spontaneous

When it is negative, the reaction is spontaneous

When delta G is zero, Here, the reaction is in equilibrium

Delta H alone cannot be used to determine the spontaneity of a reaction because (some) reactions that are exothermic at a particular temperature and may be endothermic at a different temperature.

**CHEMICAL EQUILIBRIUM**

Equilibrium is defined as the state of a system there is non-observable change in the properties of the system with respect to time. Equilibrium can either be dynamic or static. In a static equilibrium, the particles are in a stationary position e.g. a balanced see-saw

In dynamic equilibrium, the particles of the system are in constant motion. A saturated solution is a typical example of a dynamic equilibrium. The rate of dissolving is equal to the rate of precipitation in a saturated solution. In a dynamic equilibrium, physical and chemical changes usually occur.

In physical equilibrium, there is no change in the composition of the substance involved

In chemical equilibrium, there is a change of the composition of the substance involved.

For a reaction to be in equilibrium, the following must be satisfied

The reaction must be reversible

It must occur in a closed system

Factors affecting equilibrium must be kept constant

The (Gibb’s) free energy change must be zero.

A reversible reaction is one that proceeds in both directions. A reversible reaction is said to be in dynamic equilibrium if the rate of the forward reaction and the back reaction are occurring at the same rate thereby producing no net changes in the properties of the system.

LAWS OF MASS ACTION

This law states that

The rate of a chemical reaction is directly proportional to the active masses of the reactants at constant temperature and pressure

For the forward reaction,

Here, m and n are the molecularities of A and B respectively

For the backward reaction,

At dynamic equilibrium,

For partial pressure,

The relationship between

Equilibrium constant is used to express substances in gaseous state. You only consider the gaseous reactants or products and not solids

The product of the equilibrium constant for the forward and backward reaction is equal to 1

Calculate the equilibrium constant for the reverse reaction

Answer: 0.2

Find the expression for the equilibrium constant for the following equations and state the relationship between kp and kc

The only gaseous substance above is

LE CHATELIER’S PRINCIPLE

This states that if an external constraint (such as a change in temperature, pressure or concentration) is imposed on a chemical system in equilibrium, the equilibrium will shift so as to neutralize the constraint.

EFFECT OF TEMPERATURE ON EQUILIBRIUM

For an exothermic reaction: A decrease in temperature shifts the equilibrium position to the right while an increase in temperature shifts the equilibrium position to the left

For an endothermic reaction: An increase in temperature shifts the equilibrium position to the right while a decrease in temperature shifts the equilibrium position to the left

It should be noted that temperature is the only factor that will affect equilibrium constant. Generally, an increase in temperature increases the equilibrium constant while a decrease in temperature decreases the equilibrium constant.

EFFECT OF PRESSURE: Pressure affects only the gaseous reactions and there must be a change in the mole of the reactant and the product. Generally a decrease in volume corresponds to an increase in pressure which automatically changes the position of the equilibrium

EFFECT OF CONCENTRATION

When reactants are added or products are removed, equilibrium shifts to the right but when products are added or reactants removed, equilibrium shifts to the left.

EFFECT OF CATALYST ON EQUILIBRIUM POSITION

A catalyst has no effect on the position of equilibrium. The basic function of a catalyst in any reaction is to increase the rate by lowering the activation energy.

For a reaction in equilibrium, its specie is involved in the equilibrium constant expression are dissolve and gaseous species.