**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 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\_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

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:

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 and 5.4g of . 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 and 0.0089g of . 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, of a monoatomic gas (Y) combines with 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

Also, take not of the dilution principle. Dilution is the adding of water (or solvent) to reduce the concentration of the acid or base. It is also accompanied by an increase of the total volume of the whole solution

of water was added to a solution of having a concentration of 2.0 and 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 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

Solution:

Volume of = =

MM of =

Conc. Of in

Conc. In mol/dm^3 =

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