Chapter 1: Experimental Chemistry

How are physical quantities measured?

Physical quantities in chemistry

Time
-SI unit: s (seconds)
-60 seconds = 1 minute
-Measured with stopwatch

Temperature
-SI unit: K (kelvin)
-Temperature in kelvin = Temperature in celsius + 273
-Measured with thermometer

Length
-SI unit: m (metre)
-1 kilometre = 1000 metre
-Measured with measuring tape, ruler

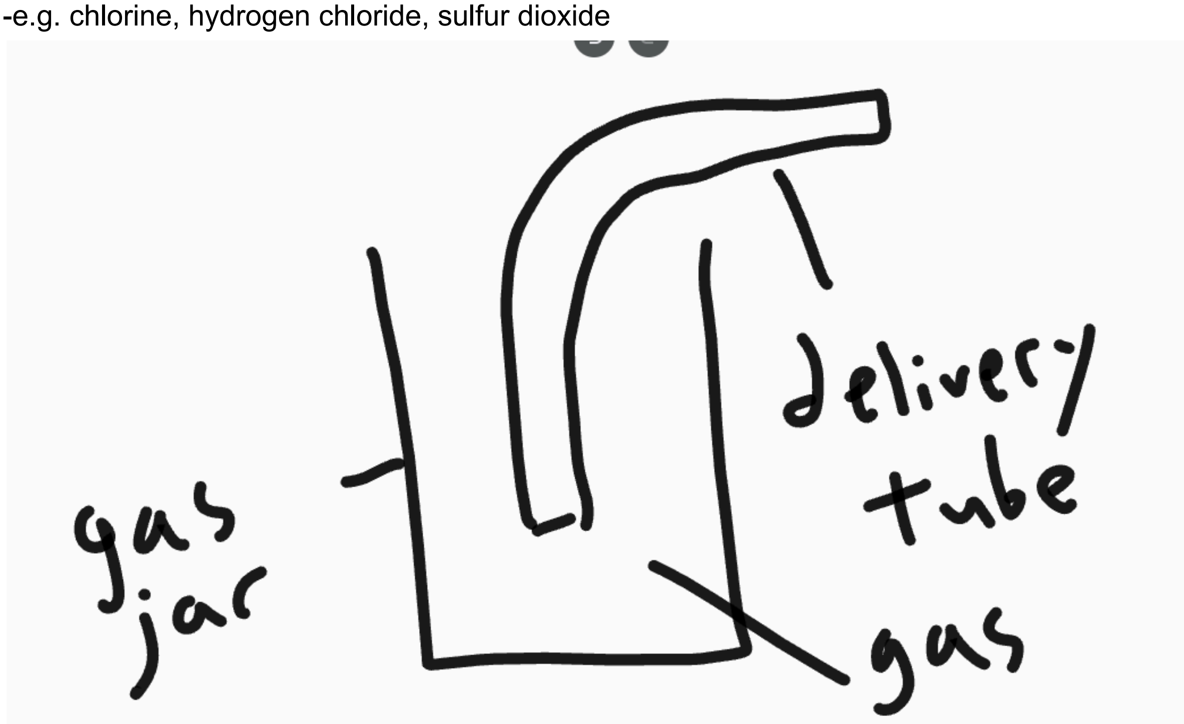
Mass
-SI unit: kg (kilogram)
-1 kg = 1000g
-Measured with beam, electronic balance

Volume
-SI unit: m^3 (cubic metre)
-1 m^3 = 100 ml
-Measured with pipette (1 d.p.), volumetric flask (big volume), measuring cylinder (approximate), burette (2 d.p.)
-Volume of gases are measured with a gas syringe

How are gases collected?

Water displacement
-Gas is bubbled through water and collected in a gas jar
-The gas can only be insoluble or slightly soluble in water
-Density does not affect this method.
-e.g. hydrogen, oxygen, carbon dioxide

Downward delivery
-Gas is collected in a gas jar and sinks down if denser than air
-Solubility in water does not affect this method
-Must be denser than air
-The Mr of air is 29
-e.g. chlorine, hydrogen chloride, sulfur dioxide



Upward delivery
-Gas is collected in a gas jar and floats up if less dense than air
-Solubility in water does not affect this method
-Must be less dense than air
-The Mr of air is 29
-e.g. ammonia

Calculating the density of gases

-Relative atomic mass (Ar): number of protons and neutrons in an atom
-The relative atomic mass of air is around 29
-Relative molecular mass (Mr): number of protons and neutrons of all atoms in a molecule
-Density = mass / volume

Methods for drying gases

Concentrated sulfuric acid
-Most gases can be dried
-Gases which react with sulfuric acid cannot be used
-The moist gas must be bubbled through the sulfuric acid

Quicklime (calcium oxide)
-Most basic and neutral gases can be used, such as ammonia
-Gases which react to calcium oxide cannot be used
-The gas must be passed through the calcium oxide
-Calcium oxide readily absorbs moisture from the air, so it must be freshly heated before use

Fused calcium chloride (fused means dried)

-Hydrogen, nitrogen, carbon dioxide is regularly used
-Gases which react with calcium chloride cannot be used
-The gas must be passed through the calcium chloride
-Calcium chloride readily absorbs moisture from the air, so it must be freshly heated before use

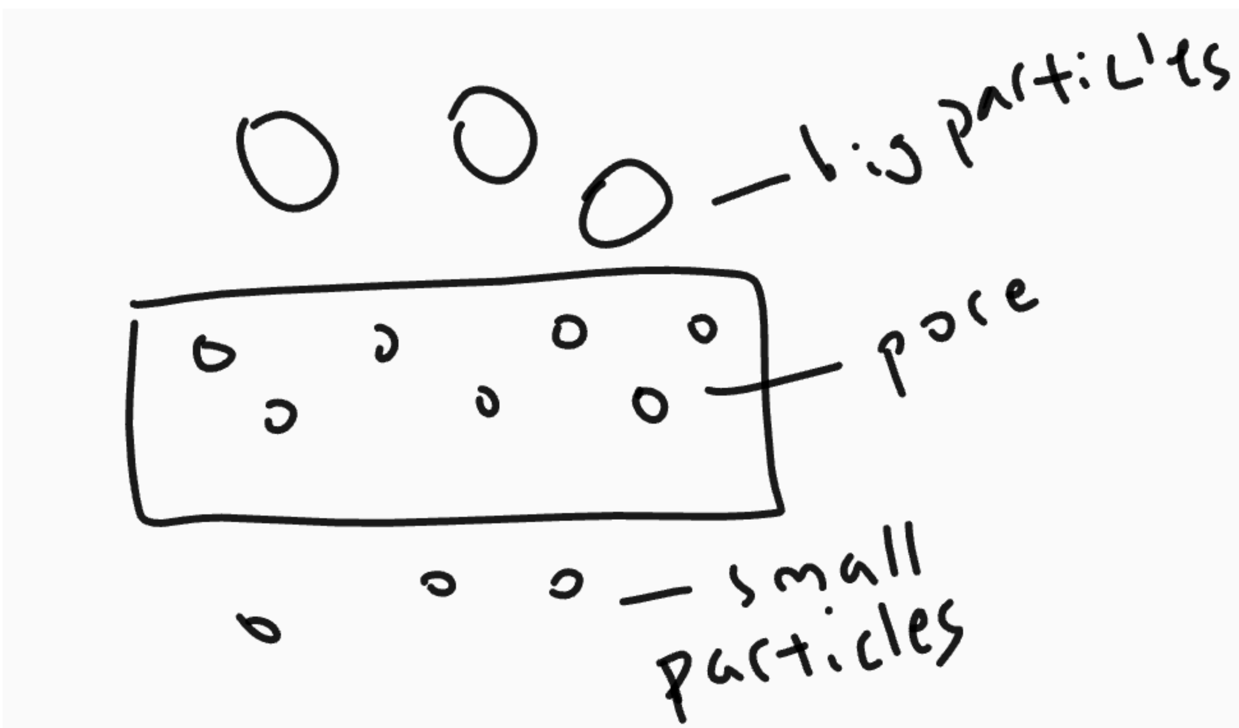
How are substances in mixtures separated?

What is a mixture?
-A mixture consists of two or more substances that are not chemically combined
-The physical properties of a substance determines the separation techniques used

Seperating solid-solid mixtures

Using magnetic attraction
-Magnetic substances can be separated from non magnetic substances
-Magnetic substances: iron, cobalt, nickel, some alloys
-Waste in a landfill is also sorted this way

Sieving
-A sieve is used to separate solids with different particle sizes
-Large particles can't pass through the smaller pores while small particles can



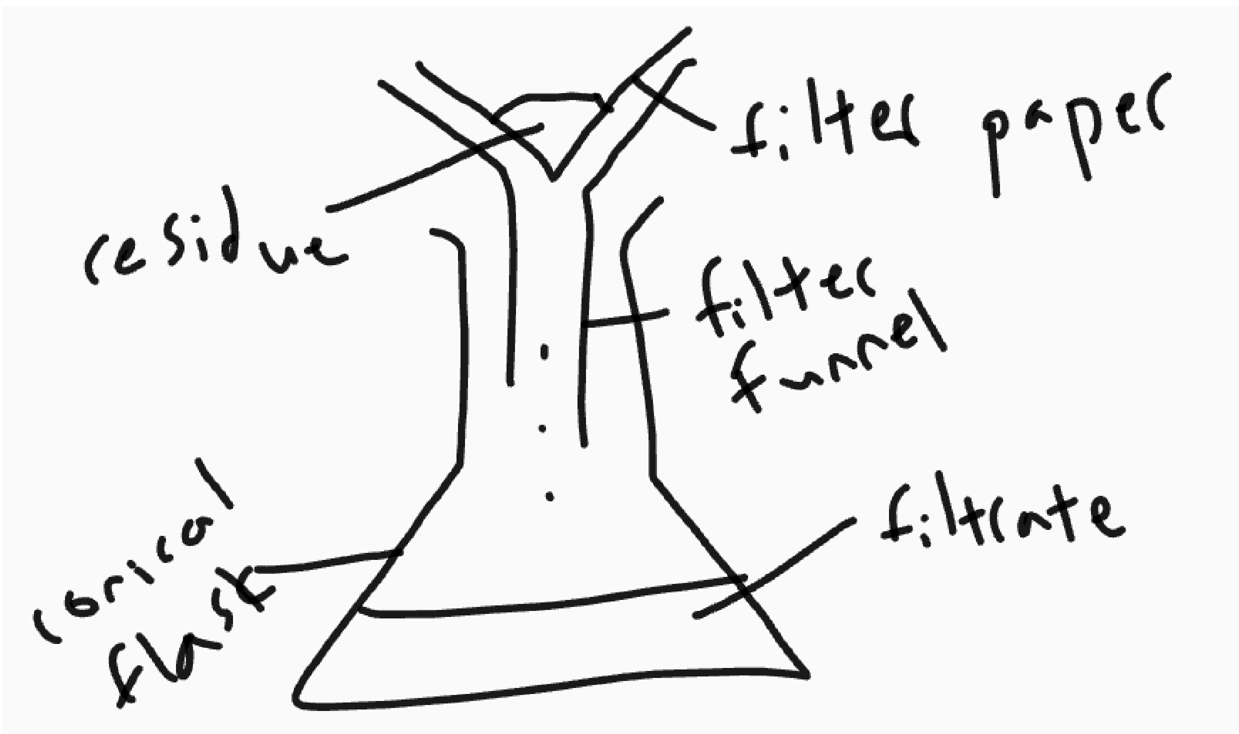
Using suitable solvents
-Solids of different solubility in the solvent can be separated
-Solvent: liquid that dissolves solids
-Solute: solid that dissolves in the solvent

-A solute dissolves in a solvent

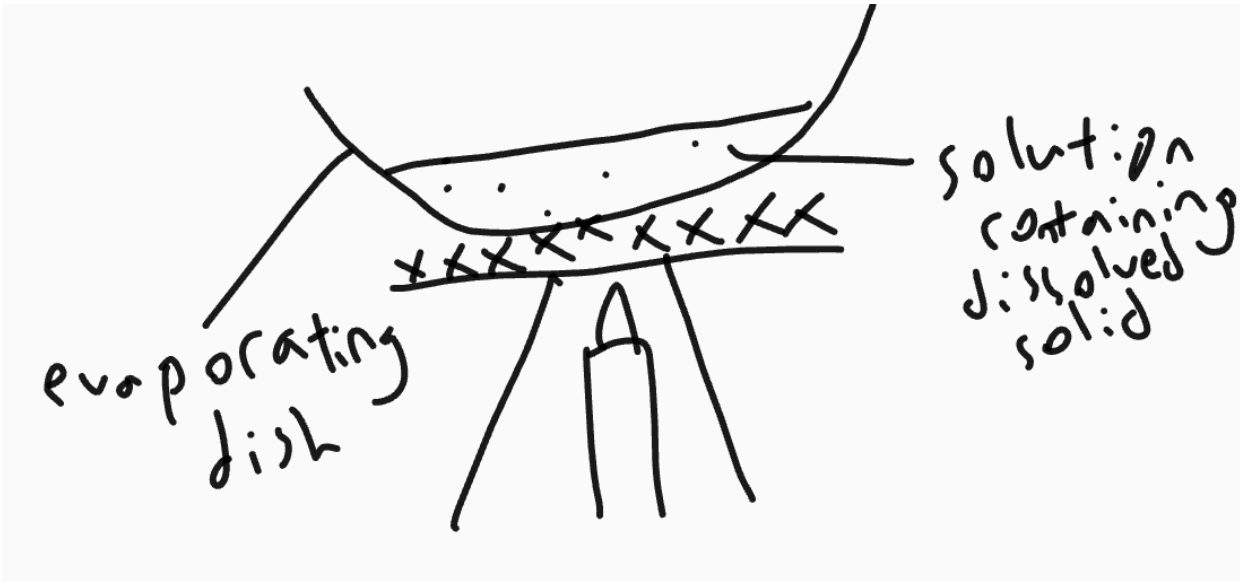
Sublimation
-Sublimation is used to separate solids where one substance changes from the solid to the gaseous state directly
-Substances that sublime: naphtalene, ammonium chloride, iodine

Separating solid-liquid mixtures

FIltration
-Seperates an insoluble solid from a liquid
-An insoluble solid is one that is unable to dissolve into the liquid it is placed in
-Filtrate: liquid that passes through the filter
-Residue: solid that remains on the filter



Evaporation to dryness
-Seperates a dissolved solid from its solvent
-The mixture is heated until all the solvent has vaporised



Crystallisation
-Seperates a pure solid from its saturated solution
-A saturated solution is one in which no more solute can be dissolved
-A solution is heated until it is saturated
-It is then cooled gradually until crystals start to form
-The crystals are then filtered from the solvent, washed with distilled water and then dried on filter paper

Simple distillation
-Seperates a pure solvent (liquid) from its solution

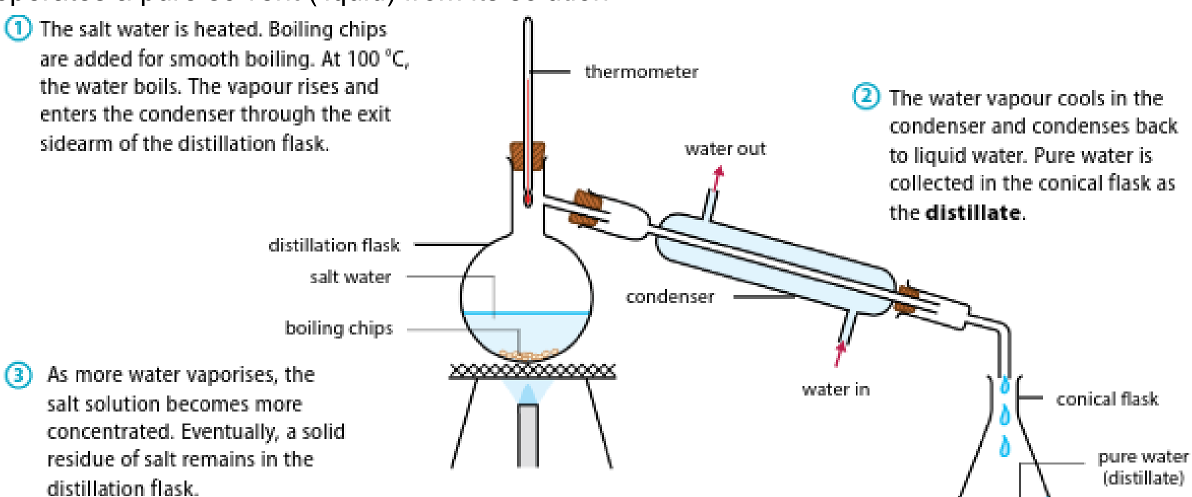


Figure 1.24 Simple distillation of salt water

-Simple distillation can only be used when there is a significant difference in boiling point

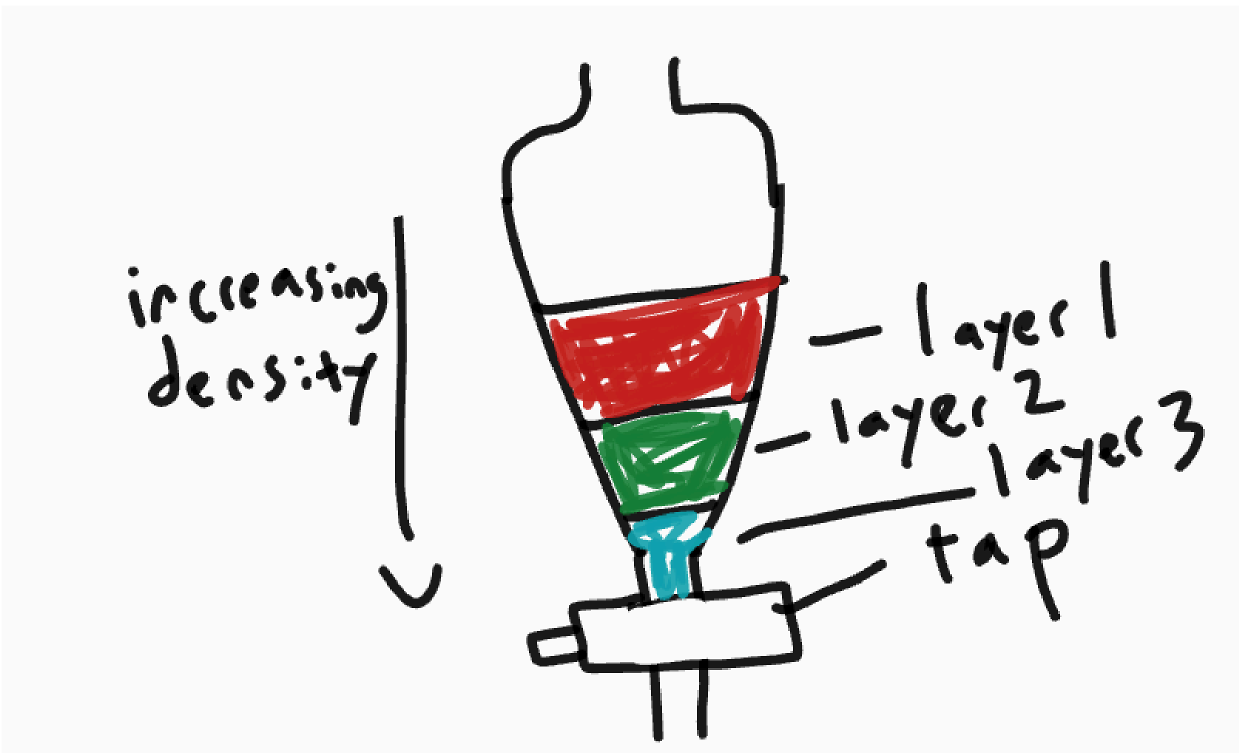
Miscible and immiscible liquids

Miscible liquids
-Miscible liquids are liquids which form a uniform (homogeneous) solution when mixed together

Immiscible liquids
-Immiscible liquids are liquids which form a non-uniform (heterogenous) mixture when mixed together

Seperating liquid-liquid mixtures

Seperating funnel
-A separating funnel is used to separate immiscible liquids
-When a heterogenous mixture is left to settle into different layers in a separating funnel, each layer can be removed by opening the tap at the bottom of the flask and collecting it in separate flasks or beakers
-The layers are separated by density, with the densest liquid at the bottom and the lest dense liquid at the top



Chromatography
-Chromatography can be used to separate miscible liquids
-A drop of a homogeneous solution is placed at the edge of a piece of paper, and indicated by a line drawn with a pencil as the start
-It is then dipped into a solvent, keeping the solvent just above the start level
-As the solvent is absorbed by the paper, the solution is separated into pure components

-They are separated by solubility in the solvent, with the most soluble at the top and the lest soluble at the bottom
-The furthest point the solvent reache is called the solvent front
-The Rf value of a substance is the ratio of the distance moved by a substance and the distance moved by the solvent until it reaches the solvent front

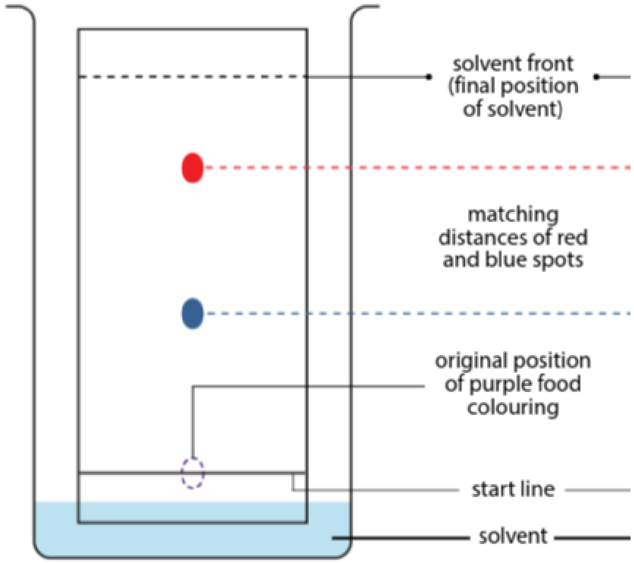


Figure 1.28 Chromatogram of a purple food colouring

Fractional distillation
-When two liquids are miscible, the separating funnel cannot be used to separate them, and the difference in boiling point of the two liquids may be too small for simple distillation to work
-Thus, a fractionating column is added to the distillation set up to separate the two liquids
-The fractionating column contains small glass beads or other small solid objects to provide a larger surface area for vapours to condense onto

Figure 1.35 shows the fractional distillation of an ethanol–water mixture.

vapour to maximise th
cooling effect.

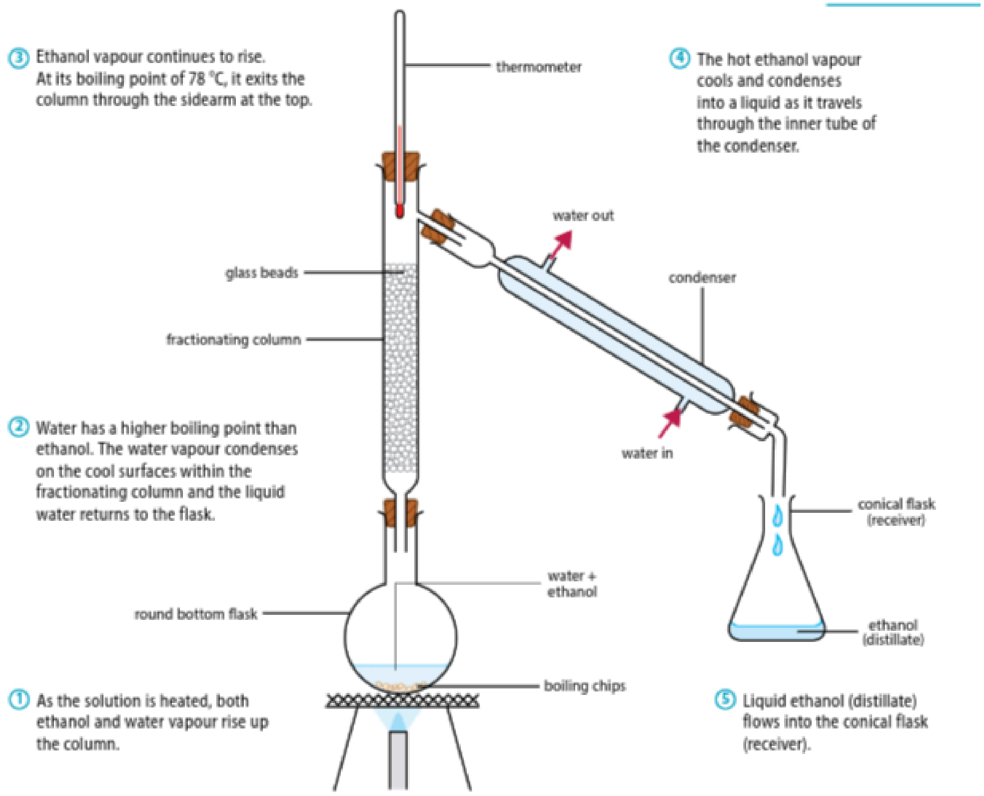


Figure 1.35 Fractional distillation of ethanol and water

How can the purity of substances be determined?

-Mixtures melt or boil at a range of temperatures, thus the greater amount of impurities in a substances, the larger the change in boiling or melting points
-Other than chromatography, melting or boiling point data can be used to determine the purity of substances

Chapter 2: Kinetic Particle Theory

Kinetic particle theory

-The kinetic particle theory states that all matter is made up of tiny particles, which are in constant random motion
-These particles are called atoms, molecules or ions

Solids, liquids and gases

-Based on the kinetic particle theory, adding or removing energy affects the forces experienced by the particles, thus determining their physical state
Solids

-Particles of a solid are very closely packed in an orderly manner, thus they only vibrate or rotate about their fixed positions and have low kinetic energy
-These particles have experience strong attractive forces, so a lot of energy is need to break up their orderly arrangement
-They have a definite volume and shape
-They undergo melting when being converted to a liquid, and undergo sublimation when converted to a gas

Liquids
-Particles of a liquid are more spaced apart compared to solids, so their attractive force is less strong
-Their arrangement is disorderly, unlike solids, and can slide past one another freely
-They have a definite volume, but no definite shape, thus taking the shape of their container
-They undergo freezing when being converted to a solid, and undergo boiling or evaporation when converted to a gas

Gases
-Particles of a gas are very spaced apart, having very weak attractive forces
-Their arrangement is disorderly, and the particles are spaced very far apart
-The particles move around quickly and randomly, and in any direction
-They have no definite volume or shape, thus taking the shape of their container and being able to be compressed
-They undergo condensation when being converted to a liquid

Changes of state

-When a substance is heated, thermal energy is transferred to the substance
-Some of the energy is then converted to kinetic energy, increasing the temperature of the substance, and vice versa when it is cooled
-At certain temperatures, heating or cooling a substances results ina change in state rather than a change in temperature
-These temperatures are known as transition temperatures
-If the temperature and time of a substances is plotted on a graph, where temperature is the y axis and time is the x axis, it is called a heating curve
-When the substance is undergoing a change of state , the temperature of the substance remains constant until all of the substance has changed states

Expansion and contraction

-When a substance is heated below its melting point, the particles cannot spread out freely due to the strong attractive force
-Thus, they vibrate more quickly about their positions, but with a slightly wider spacing than before, leading to expansion
-When a substance is cooled the opposite happens and the particles come closer than before, leading to contraction

Movement of particles

-In liquids, the dissolved particles of a substance will move in the liquid through diffusion
-In gases, the solid particles will collide with the gas particles and move around through diffusion
-Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration (bio definition with concentration gradient is accepted in chem exam, but chem definition is not accepted in bio exam)

Factors affecting diffusion

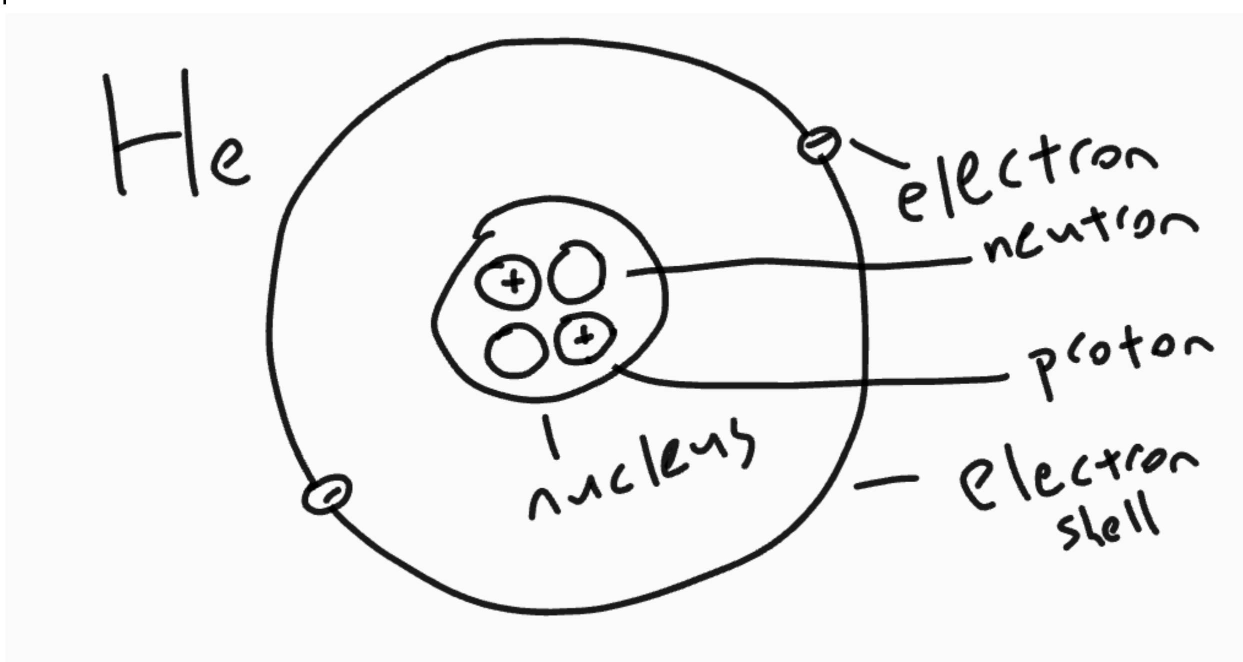
Temperature
-If a particle has a higher temperature, it has more kinetic energy, thus moving around more quickly
-Thus, the higher the temperature, the faster the rate of diffusion

Molecular or atomic mass
-Particles with a larger mass require more kinetic energy to move at a given speed
-Thus, the higher the molecular or atomic mass, the slower the rate of diffusion

Chapter 3: Atomic Structure

What is an atom made up of?

-An atom is the smallest particle that can have the chemical characteristics of an element
-However, there are even smaller particles inside the atom called sub atomic particles
-Different elements have different properties as they have different amounts of sub atomic particles



Sub atomic particles

-Atoms are electrically neutral as the charges of the proton and electron cancel out
-Number of protons = number of electrons
-Protons and nucleus are found in the nucleus, and are thus called nucleons
-An atom has affixed amount of protons, however it can have a different number of neutrons to form different isotopes
-If an atom gains an electron, it becomes a negative ion (anion)
-If an atom loses an electron, it becomes a positive ion (cation)

Sub atomic particle	Relative mass	Relative charge	Location in an atom
proton	1	+1	nucleus
neutron	1	0	nucleus
electron	1/1840 / negligible	-1	electron shell

How many sub atomic particles does an atom have?

Proton (atomic) number (Z)
-Number of protons

Nucleon (mass) number (A)
-Number of protons + number of neutrons

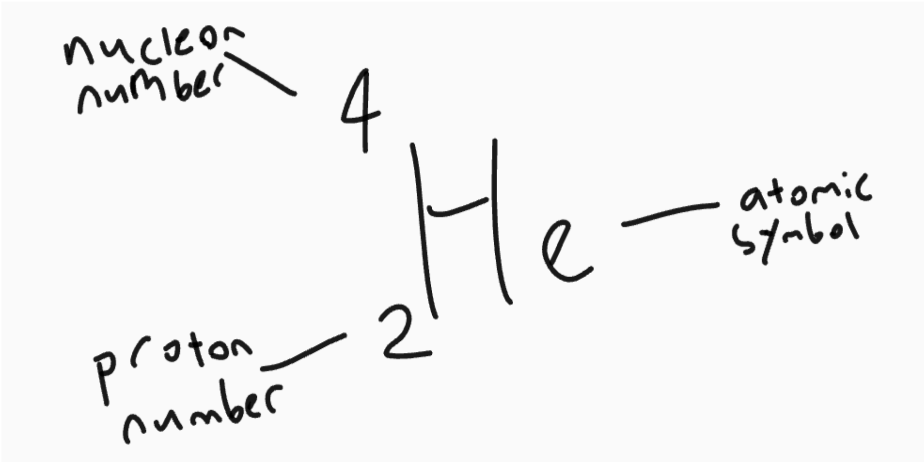
Neutron number
-Nucleon number - proton number

Ar
-Ar is the relative atomic mass of an atom
-It is the nucleon number

Mr
-Mr is the relative molecular mass of an atom
-It is the total nucleon number of all atoms in a molecule (2 or more atoms chemically bonded together)

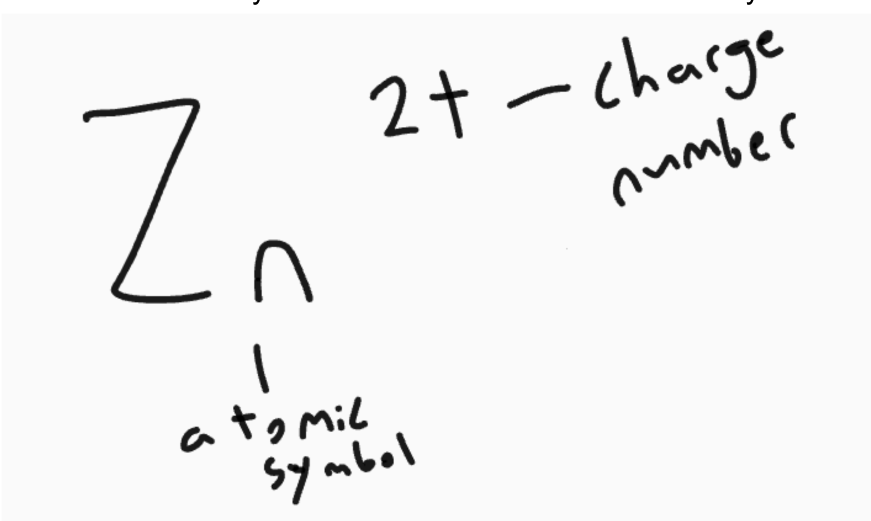
Nuclide notation

-Nuclide notation shows the proton number and nucleon number



Ions

-An ion is the particle formed by the replacement of a hydrogen atom in an acid
-We use charge numbers to indicate the number of electrons that are gained or lost
-A positive ion is called a cation, while a negative ion is called an anion
-Metals usually lose electrons while non metals usually gain electrons
-Thus metals usually form cations while non metals usually form anions



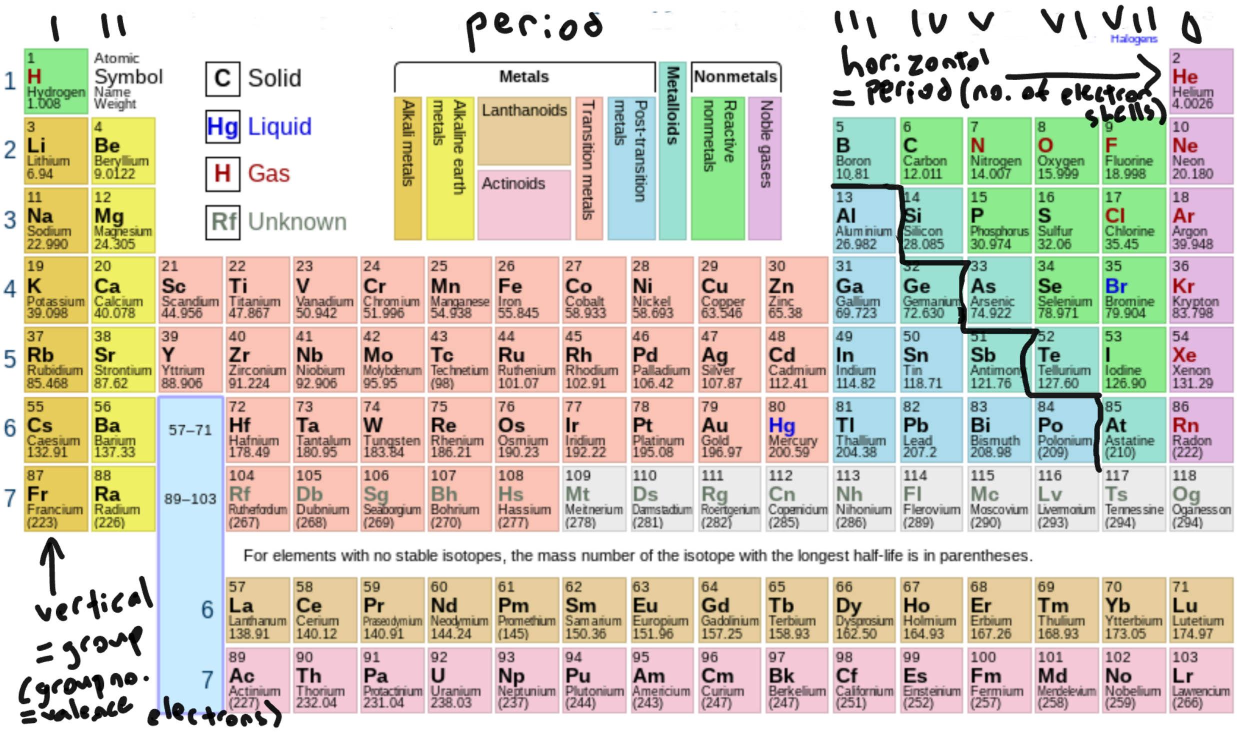
Isotopes

-Isotopes are atoms of the same elements that have the same proton number but different nucleon numbers
-Since the proton number is the same, the isotopes belong to the same element
-Most elements occur naturally as isotopes
-The nucleon number used in the periodic table is found by finding the relative abundance of all isotopes of an element on earth
-e.g. out of all chlorine on earth, 75% is chlorine-35 and 25% is chlorine-37, thus [75(35) + 25(37)] / 100 = 35.5 (nucleon number on periodic table)

-Formula for relative abundance: [p1(m1) + p2(m2)] / 100, where p is percentage on earth and m is nucleon number

-They have the same chemical properties, but different physical properties as they have the same amount of protons but different amounts of neutrons

Reading the periodic table



(in exam round nucleon number to nearest whole number, except for chlorine which is 35.5 and take all elements after bismuth as having no nucleon number on the periodic table)

How are sub atomic particles distributed in an atom?

-Electrons are constantly moving around the nucleus in zones called electron shells

-Electrons in the outermost shell (valence shell) possess the most energy

-Inner electron shells are filled first

-The first electron shell can hold 2 electrons, and the second and third can usually hold 8 electrons each

-When the outermost shell is fully filled, the atom is stable and unreactive (inert)

-The electrons in the valence shell are known as valence electrons

Electronic configuration

-Electronic configuration is the way you write the amount of electrons in an atom

-e.g. the electronic configuration of oxygen is 2.6 as there are 2 electrons in the first shell and 6 in the second shell

Chapter 4: Chemical Bonding

Why do atoms combine?

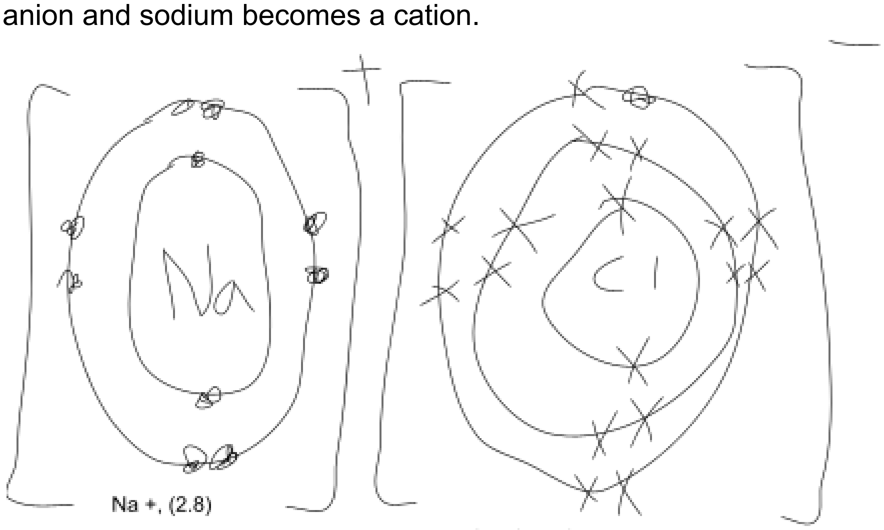
-Some atoms which do not have full valence electron shells combine to achieve the stable electronic configuration of a noble gas
-This occurs through the sharing, gain or loss of electrons

Noble gases

-Noble gases occupy the last group of the periodic table
-They are considered chemically unreactive as they have a full valence electron shell
-They are monoatomic

Ionic bonding

-When a metallic atom reacts with a non metallic atom, the metallic atoms give their electrons to the non metallic atom. This forms two ions which both have the electronic configuration of a noble gas.
-When an atom loses an electron, it becomes a positive cation. Likewise, if an atom gains an electron, it becomes a negative anion.
-If an atom loses two electrons, it becomes a “2+” cation and vice versa.
-If the metallic atoms has not enough electrons to give, then add more of that atom.
-Chlorine has an electronic configuration of 2.8.7. The nearest noble gas has an electronic configuration of 2.8.8. Thus, it needs one electron to attain that electronic configuration. When reacting with sodium, which has an electronic configuration of 2.8.1, it loses one electron to chlorine, attaining the electronic configuration of a noble gas, or 2.8. Chlorine then becomes an anion and sodium becomes a cation.



Na +, (2.8)

Cl -, (2.8.8)

Common cations

Relative charge	Name of cation	Formula
+1	hydrogen	H^+
+1	sodium	Na^+
+1	potassium	K^+
+1	silver	Ag^+
+1	all group 1 alkali metals and sometimes hydrogen	nil
+1	ammonium	NH_4^+ (there are 4 hydrogen atoms, not a charge of +4)
+2	magnesium	Mg^{2+}
+2	calcium	Ca^{2+}
+2	barium	Ba^{2+}
+2	iron (II)	Fe^{2+}
+2	copper (II)	Cu^{2+}
+2	zinc	Zn^{2+}
+2	lead (II)	Pb^{2+}
+2	all group 2 alkaline earth metals	nil
+3	iron (III)	Fe^{3+}
+3	all group 13 elements	nil
+3	aluminium	Al^{3+}
+4	some group 14 elements	nil

Common anions

Relative charge	Name of anion	Formula
-1	fluoride	F^-

-1	chloride	Cl-
-1	bromide	Br-
-1	iodide	I-
-1	hydroxide	OH-
-1	nitrate	NO3- (there are 3 oxygen atoms, not a charge of -3)
-1	manganate (VIII)	MnO4- (there are 4 oxygen atoms, not a charge of -4)
-1	all group 17 halogens and sometimes hydrogen	nil
-2	oxide	O2-
-2	carbonate	CO32- (there are 3 oxygen atoms and a charge of -2, not a charge of -32)
-2	sulfate	SO42- (there are 4 oxygen atoms and a charge of -2, not a charge of -42)
-2	all group 16 elements	nil
-3	phosphate	PO43- (there are 4 oxygen atoms and a charge of -3, not a charge of -43)
-3	all group 15 elements	nil
-4	some group 14 elements	nil

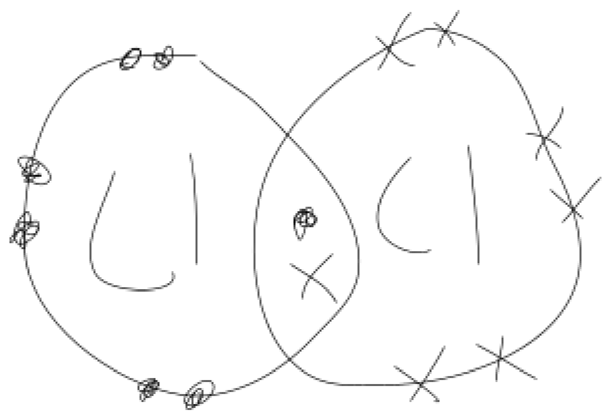
(note that in exam, the number for amount of atoms is written as a subscript while charge is a superscript)

Ionic compounds

-Ionic compounds are two ions held together by ionic bonds, or the strong electrostatic forces between positive and negative ions
-Most ionic compounds are formed by at least one metal cation and one non metal anion
-They have no net charge as the charges cancel out
-Like all compounds, they are formed in a fixed ratio

Covalent bonding

-Non metallic atoms, when taking part in chemical reactions,will try to attain the electronic configuration of noble gases by sharing electrons.
-The number of electrons shared is determined by the number of valence electrons, and the number shared is the same.
-Chlorine has an electronic configuration of 2.8.7. The nearest noble gas has an electronic configuration of 2.8.8. Thus, if two chlorine atoms were to take part in covalent bonding, they would each share on electron so that both atoms attain the electronic configuration of 2.8.8.



Covalent compounds

-Covalent compounds are two or more atoms or molecules held together by either weak intermolecular forces of attraction or strong covalent bonds
-Most covalent compounds do not contain metal atoms
-They have no net charge as the charges cancel out
-Like all compounds, they are formed in a fixed ratio

Chapter 5: Structure and Properties of Materials

Elements

-An element is a pure substance which cannot be broken down into simpler substances by chemical methods
-They can appear as atoms when not bonded to other atoms, molecules when bonded to another of the same atom, or a lattice in a giant metallic structure
-They have a fixed melting and boiling point

Compounds

-A compound is a pure substance which consists of two or more elements chemically combined in a fixed ratio
-The elements can bond through ionic or covalent bonding

-Not all molecules are compounds, as hydrogen gas (H_2) is an element which exists as a molecule
, while carbon dioxide (CO_2) is a compound
-They can be separated, although with difficulty, through processes such as heat decomposition or electrolysis
-They are formed from a chemical reaction and have different properties as compared to their constituent elements
-They have a fixed melting and boiling point

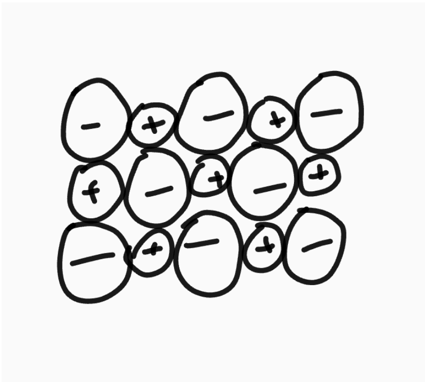
Mixtures

-A mixture is an impure substances consisting of elements, molecules or compounds physically combined together without a fixed ratio
-For example, air is a mixture as it contains various gases, such as hydrogen, oxygen, nitrogen and carbon dioxide
-The gases in air are not combined in a fixed ratio, as air at higher elevations may have different ratio of gases
-Generally, they can be easily separated through heating or filtration
-However, alloys require a large amount of energy to separate
-They are usually formed from physical mixing and have similar properties to their constituent elements, molecules or compounds
-They have a range of melting and boiling points

Giant Ionic lattice structures (GILS)

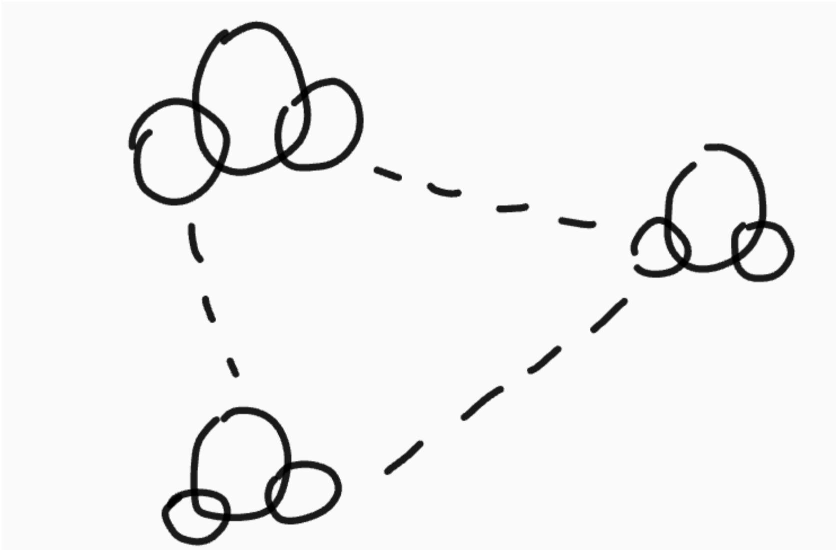
-In the solid state, ionic compounds exist as a 3D structure known as a giant ionic lattice structure
-Each ion is placed next to an ion with the opposite charge
-Thus, each ion is surrounded by 6 ions of an opposite charge
-They have a regular shape as the ions are arranged in an orderly and regular manner
-They are very hard but brittle as the strong electrostatic forces of attraction hold them together
-However, once these bonds are broken, the ions move away from each other and atoms of the same charge move towards each other, increasing the repulsive force between the ions and eventually shattering the structure when the repulsive forces are stronger than the attractive forces
-They have a high melting and boiling point as a lot of heat energy is needed to overcome the strong electrostatic forces of attraction holding them together
-When a giant ionic lattice structure is made up of doubly charged ions (+2, -2), they have a higher melting and boiling point as compared to singly charged ions
-They are soluble in water as the water molecules are attracted to the ions, weakening the electrostatic forces between the ions
-The ions are thus pulled from the lattice structure, dissolving to form an aqueous solution
-They are not soluble in organic solvents

-They can conduct electricity in molten and aqueous states but not solid states, as the rigid lattice structure is broken down and the charged ions are allowed to move around and carry charges, while in solid state, the ions are not able to move around



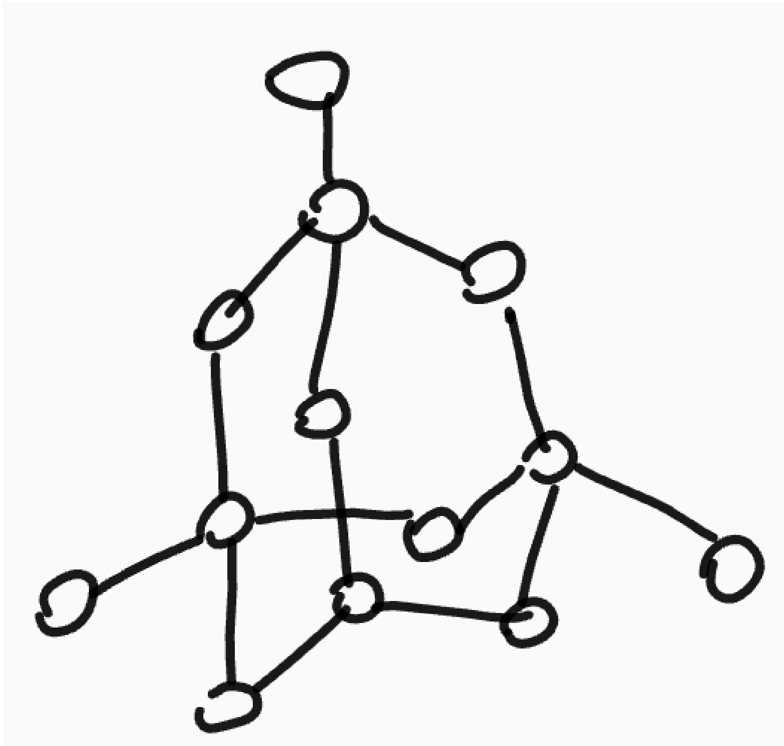
Simple molecular substances (SMS)

-SImple molecular substances, or simple covalent substances, are formed by covalent molecules
-There are weak intermolecular forces between the molecules in the structure of the simple molecular substance
-Thus, they have low melting and boiling points as little heat energy is needed to overcome these forces and are quite volatile (melting and boiling point close to room temperature)
-The larger the size of the molecules, which is proportional to the number of electron shells, the stronger the intermolecular forces of attraction and the higher the melting and boiling points
-They are insulators of electricity as each valence electrons of an atom are used in covalent bonding, so there are no mobile charge carriers
-However, acids which are covalent compounds, such as hydrogen chloride and hydrogen fluoride, can dissolve in water and dissociate into their respective ions, will be able to carry charges and conduct electricity
-However, these compounds are not ionic



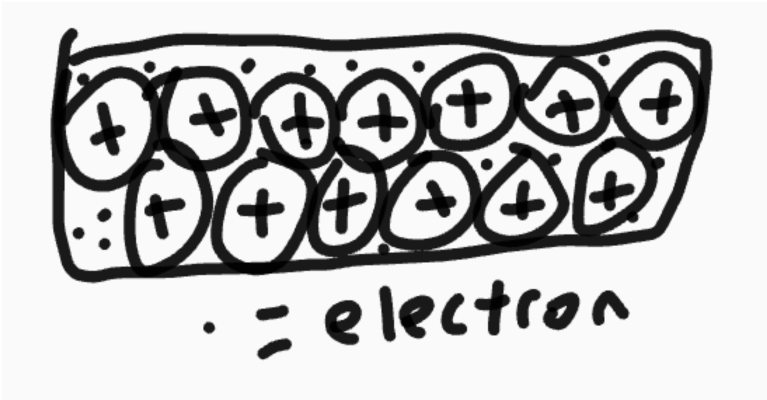
Giant molecular substances (GMS)

-Giant molecular substances, or giant covalent substances, are formed by covalent molecules or individual atoms
-They are joined by strong covalent bonds, unlike simple molecular substances
-Common giant molecular substances include diamond, graphite, silicon and silicon dioxide
-They have a high melting and boiling point as a large amount of energy is needed to break the strong covalent bonds
-They are insoluble in water and organic solvents as the forces of attraction between the solvent molecules and the atoms of the giant molecular substance are not strong enough to break the strong covalent bonds
-They are insulators of electricity as each valence electrons of an atom are used in covalent bonding, so there are no mobile charge carriers
-They are very hard due to the strong covalent bonds between the atoms
-However, graphite can conduct electricity as only 3 of the 4 valence electrons are used for bonding, and the extra valence electron can move and carry a charge
-Graphite is also quite soft as graphite is made up of different layers of carbon atoms, joined by weak intermolecular forces of attraction, which can be overcome easily when a force is applied



Giant metallic substances

-Giant metallic substances are described as a lattice of metal cations, surrounded by a sea of delocalised valence electrons
-These electrons can move and carry a charge, so metals have high electrical and thermal conductivity
-Metals also have a high melting and boiling point, are malleable and ductile, are sonorous (makes a ringing sound when striked), are shiny, are dense, form alkaline oxides and react by losing electrons to form cations
-They have high melting and boiling points as a lot of energy is needed to overcome the strong electrostatic forces of attraction between the metal cations and surrounding sea of delocalised valence electrons
-They are also ductile and malleable as the atoms of pure metals are of the same size and packed in orderly layers, allowing the layers of atoms to slide over each other easily when force is applied



Alloys

-Alloys are a mixture of a metal with another element
-Brass (copper and zinc) and steel (iron and carbon) are both alloys
-They are usually stronger and harder, have increased corrosive resistance and have a lowered melting point (when an impurity is added to a substance, the melting point decreases while the boiling point increases) compared to pure metals, thus they are more commonly used compared to pure metals
-Alloys are strong and hard due to the different sizes of atoms disrupting the orderly layers of atoms and making it difficult for the layers to slide over each other when a force is applied

Chapter 6: Chemical Formulae and Equations

Anions

-Anions are negatively charged non metal ions
-They gain electrons when undergoing ionic bonding with a cation
-When writing the name of the anion in a ionic compound, it has the name of the base element but replaced with -ide at the end
-For example, an ion of fluorine is known as a fluoride ion

Cations

-Cations are positively charged ions, usually made from metals, but not always
-They lose electrons when undergoing ionic bonding with a anion
-Most transition metals (group 3-12) and post transition metals (group 13-15 metals) form a cation with more than one possible charge, which is depicted in brackets with the charge in roman numerals
-For example, a Cu2+ cation is written as copper (II)
-However, some transition and post transition metals only form cations with one charge, such as silver and zinc

-These are written by their names only

Polyatomic ions

-Polyatomic ions are formed from molecules which have gained or lost electrons
-Negatively charged polyatomic ions are also written with an -ide at the end
-When writing more than one polyatomic ion bonded togther, it is written with brackets
-For example, calcium hydroxide is Ca(OH)2
-A compound that contains a polyatomic ion containing oxygen ends with an -ate
-Almost all polyatomic ions are negatively charged

Group number relative to charge of ion

Group number	Charge of ion
1	+
2	+2
13	+3
14	+4 / -4
15	-3
16	-2
17	-

Common formulae

Name	Formula
hydrochloric acid	HCl
sulfuric acid	H2SO4
nitric acid	HNO3
sodium hydroxide	NaOH
potassium hydroxide	KOH
calcium hydroxide	Ca(OH)2
hydroxide ion	OH-
water	H2O

ammonia (not to be confused with ammonium)	NH_3
methane	CH_4
sulfate ion	SO_4^{2-} (there are 4 oxygen atoms and a charge of 2-, not a charge of 42-)
carbonate ion	CO_3^{2-} (there are 3 oxygen atoms and a charge of 2-, not a charge of 32-)
nitrate ion	NO_3^- (there are 3 oxygen atoms, not a charge of 3-)
phosphate ion	PO_4^{3-} (there are 4 oxygen atoms and a charge of 3-, not a charge of 43-)
ammonium ion	NH_4^+ (there are 4 hydrogen atoms, not a charge of 4+)
zinc ion	Zn^{2+}
silver ion	Ag^+
copper (II) ion	Cu^{2+}
lead (II) ion	Pb^{2+}
iron (II) ion	Fe^{2+}
iron (III) ion	Fe^{3+}

Writing ionic compounds

-Write down the charges of the cation and the ion
-Add cation or anions in order for the charges to cancel out
-Always write the cation before the anion
-For example, calcium has a charge of 2+ and bromine has a charge of -
-So, you need 2 bromine atoms to cancel out the 2+ charge
-They then form $CaBr_2$, or calcium bromide
-This also works for polyatomic ions
-For example, zinc has a charge of 2+ and a nitrate ion has a charge of -
-So, you need 2 nitrate ions to cancel out the 2+ charge
-They then form $Zn(NO_3)_2$, or zinc nitrate

Writing covalent compounds

-Writing covalent compounds is quite similar to ionic compounds, except they need a prefix

Prefix	Number of atoms
mono-	1
di-	2
tri-	3
tetra-	4

-For example, for BCl_3, there is 1 boron atom and 3 chlorine atoms
-Thus, it is written as boron trichloride (there is no mono prefix for boron as it is the first element)

Chemical equations

-A chemical equation is an equation which shows a chemical reaction and the reactants and products
-The reactants are always written on the left while the products are on the right
-The equation is written with an arrow instead of the equal sign
-There are state symbols written next to the chemicals showing the state they are in
-Common state symbols are s (solid), l (liquid), g (gas) and aq (aqueous)
-Aqueous means a solid dissolved in a water to form a solution
-For example, when reacting hydrochloric acid and sodium hydroxide solution, it is written as HCl (aq) + $NaOH$ (aq) -> $NaCl$ (aq) + H_2O (l)

Balancing chemical equations

-When writing chemical equations, the atoms on each side of the equation must be balanced, meaning that there must be the same amount of each atom
-This can be done by writing the unbalanced chemical equation down, then writing the amount of each atom on each side
-If the number of atoms do not match, add molecules of the unbalanced atoms to the left side until they are balanced
-Only add molecules and not singular atoms
-For example, the equation Na (s) + Cl_2 (g) -> $NaCl$ (s) is not balanced
-There is 1 sodium atom and 2 chlorine atoms on the left side, while there is 1 sodium and chlorine atom on the right side
-Thus, 1 sodium atom needs to be added to the left side
-The balanced equation is then written as $2Na$ (s) + Cl_2 (g) -> $2NaCl$ (s)

Chapter 7: Mole Concept and Stoichiometry

Relative atomic mass

-The relative atomic mass (Ar) is the average mass of one atom of an element relative to 1/12 the mass of an atom of carbon-12

Relative molecular mass

-The relative molecular mass (Mr) of a molecular substance is the average mass of one molecule of that substance relative to 1/12 of the mass of an atom of carbon-12
Percentage mass
-Percentage mass is the percentage of a certain element in a compound
-Formula for percentage mass: [(number of atoms of the element x Ar of the element) / Mr of compound] x 100%

What is a mole?

-A mole is the amount of substance which contains 6.02 x 10^23 particles
-SI unit: mol (mole)
-A particle can be atoms, ions, molecules or even sub atomic particles

Avogadro number

-The avogadro number or avogadro constant is defined as the number of atoms in a 12 g sample of carbon-12
-The avogadro number is equal to 6.02 x 10^23

Moles and mass

-The mass of an atom or molecule is equal to its relative atomic mass or relative molecular mass in grams
-Formula for number of moles: mass / Ar or Mr
-Formula for mass: number of moles x Ar or Mr

Molar gas volume

-One mole of any gas at room temperature and pressure (r.t.p) occupies a volume of 24000 cm^3
-Molar gas volume at r.t.p is 24000 cm^3
-Thus, equal volumes of all gases at r.t.p contain the same number of particles
-Formula for volume of any gas in cm^3: number of moles x molar gas volume in cm^3

Empirical formula

-The empirical formula of any compound has the general formula AxBy
-To find the empirical formula, we find the number of moles of each atom in the formula, then find the ratios of each element in the formula
-For example, 9.6 g of carbon, 25.6 g of oxygen and 0.8 g of hydrogen are in a compound
-No. of mol of C = 9.6 / 12 = 0.8 mol
-No. of mol of O = 25.6 / 16 = 1.6 mol
-No. of mol of H = 0.8 / 1 = 0.8 mol
-Ratio of C:O:H = 0.8:1.6:0.8 = 1:2:1

-Thus the compound is CHO_2
-When writing the formula, carbon is written first, followed by hydrogen and the rest of the elements in alphabetic order of chemical symbols
-If carbon is not present, all elements including hydrogen are written in alphabetic order of chemical symbols
-Do note that for gases, do not use the molecular mass of each element if they are diatomic, always use the atomic mass instead

Molecular formula

-The molecular formula of any compound has the general formula $n(A_xB_y)$, where n is any number
-Formula for n: Mr of molecule / Mr from empirical formula
-For example, the empirical formula of propane is CH_2, and the relative molecular mass of propane is 42
-Mr of empirical formula = 12 + 2 = 14
-n = 42 / 14 = 3
-Thus, the molecular formula of CH_2 = $(CH_2)_3$ = C_3H_6

Moles and equations

-In chemical equations, the number of molecules of a reactant is equal to molar ratio of the reactants
-For example, in the reaction $2H_2 + O_2 = 2H_2O$, there are 2 molecules of hydrogen gas and one molecule of oxygen gas, so the molar ratio of hydrogen:oxygen:water is 2:1:2
-Using this, we can find the mass of a product given a fixed amount of reactant
-For example, if there are 10 g of hydrogen gas, we can find the mass of water formed
-Number of moles of H_2 = 10 / 2 = 5
-Mole ratio of H_2:H_2O = 2:2
-Number of moles of H_2O = 5
-Mass of H_2O = 5 x 18 = 90 g

Percentage purity

-Most chemicals used in a chemical reaction are almost never 100% pure, as they will have some impurities in them
-Formula for percentage purity of a substance: (mass of pure substance used in reaction / mass of impure substance) x 100%
-For example, if 12 g of impure carbon is burned in the presence of oxygen to form 30 g of pure carbon dioxide in the reaction$C + O_2 = CO_2$, we can find the percentage purity of the carbon
-Number of moles of CO_2 formed = 30 / 44 = 0.682 mol (3s.f.)
-Actual number of moles of C reacted = 0.682 mol (molar ratio of C:CO_2 is 1:1)
-Actual mass of C reacted = 0.682 x 12 = 8.18 g (3s.f.)
-Percentage purity of C = (8.18 / 12) x 100% = 68.2% (3s.f.)

Percentage yield

-In a chemical reaction, the actual amount of product will always be less than the theoretical amount of product
-Formula for percentage yield of a substance: (actual yield / theoretical yield) x 100%
-For example, 1.92 g of magnesium metal was heated in the presence of oxygen to form 3 g of magnesium oxide in the reaction $2Mg = O_2 = 2MgO$
-Number of moles of Mg = 1.92 / 24 = 0.08 mol
-Number of moles of MgO = 0.08 mol (molar ratio of Mg:MgO is 2:2)
-Theoretical yield of MgO = 0.08 x 40 = 3.2 g
-Percentage yield = (3 / 3.2) x 100% = 93.8% (3s.f.)

Moles and solution

-When a soluble substance is dissolved in a liquid the substance is known as the solute, the liquid is known as the sovent and the resultant liquid is called a solution
-In general, we usually use water as the solvent in this chapter, and thus the solution is aqueous
-The concentration of a solution is recorded in mol/dm^3 or g/dm^3
-1 dm^3 = 1000 cm^3 = 1000 ml
-Formula for concentration in mol/dm^3 = no. of moles of substance / volume of solution in dm^3
-Formula for concentration in g/dm^3 = concentration in mol/dm^3 x Mr of solute
-If an acid is monobasic, 1 mol/dm^3 of acid will contain 1 mol/dm^3 of H^+ ions, and 1 mol/dm^3 of a dibasic acid will contain 2 mol/dm^3 of H^+ ions and so on

Limiting reactant

-When doing chemical reactions, chemists will limit the amount of a certain reactant to ensure it stops at a certain point
-The reactant with fewer moles is the limiting reactant, and the reactant with more moles is the excess reactant
-For example in the reaction Zn + S = ZnS, where 25 g of zinc is used and 30 g of sulfur is used, we can find the limiting reactant
-Number of moles of ZnS is all Zn is used = 25 / 65 = 0.385 mol (3s.f.)
-Number of moles of ZnS if all S is used = 30 / 32 = 0.938 mol (3s.f.)
-Since zinc forms fewer moles of zinc sulfide, zinc is the limiting reactant and sulfur is the excess reactant

Moles and titration

-When conducting a titration, we can find an unknown value in the titration using this formula: [(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of acid = [(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of alkali
-For example, if 30 cm^3 of sodium hydroxide is used to neutralise 25 cm^3 of 0.1 mol/dm^3 of sulfuric acid in the formula $2NaOH + H_2SO_4 = Na_2SO_4 + 2H_2O$, we can determine the concentration of the sodium hydroxide using the formula above
-By plugging the values into the formula, we get [(0.1 x 25) / 1] = [(M x 30) / 2]

-When solving for M, which is concentration, we get 0.167 mol/dm^3 (3s.f.)

Summary of formulas

-No. of moles (solid, liquid) = mass in grams / Ar or Mr
-Mass in grams (solid, liquid) = no. of moles x Ar or Mr
-No. of particles = no. of moles x 6.02 x 10^23
-Volume of gas in dm^3 = no. of moles x 24
-Volume of gas in dm^3 = (mass in grams / Mr) x 24
-No. of moles (gas) = volume in dm^3 / 24
-Mass in grams (gas) = (volume in dm^3 / 24) x Mr
-Empirical formula = AxBy
-Molecular formula = n(AxBy)
-n = Mr of molecule / Mr from empirical formula
-Percentage purity = (mass of pure substance used in reaction / mass of impure substance) x 100%
-Percentage yield = (actual yield / theoretical yield) x 100%
-Concentration in mol/dm^3 = no. of moles of substance / volume of solution in dm^3
-Concentration in g/dm^3 = concentration in mol/dm^3 x Mr of solute
-[(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of acid = [(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of alkali

Chapter 8: Acids and Bases

What is an acid?

-An acid is a substance which dissociates/ionises in water to form H+ or hydrogen ions
-They have a sour taste and are generally corrosive
-they are electrolytes, which means they can conduct electricity in aqueous solutions
-They turn moist litmus paper red and turn green universal indicator orange if weak or red if strong
-They have a pH below 7

Basicity of acids

-The basicity of an acid is the maximum number of H+ ions produced per mole of acid
-A monobasic acid will produce 1 H+ ion, a dibasic acid will prduce 2, a tribasic acid will produce 3 and so on

Mineral vs organic acids

-A mineral acid is an acid which does not occur naturally and is usually man made
-Examples include sulfuric, hydrochloric and nitric acid
-These acids are strong, corrosive and for laboratory and industrial use
-Organic acids occur naturally in fruits and plants

-Examples include citric, ethanoic and tartaric acid
-They are weak, less corrosive and can be used for flavouring food and cooking

Strength vs concentration

-The strength of an acid refers to the extent to which an acid molecule dissociates to forms H+ ions
-A strong acid will dissociate completely to form a high concentration of H+ ions, while a weak acid will dissociate partially to form a low concentration of H+ ions
-The strength of an acid cannot be changed
-However, the concentration of a solution refers to the amount of solute dissolved into 1 dm^3 of water
-A solution can be either dilute or concentrated
-Concentration can be changed by adding more or less solute

Test for acids

Test	Result
2-3 drops of universal indicator	universal indicator turns from green to orange or red depending on the strength of the acid
2-3 drops of methyl oange	methyl orange turns from orange to red
add 2 g of calcium carbonate to the acid	effervescence observed, and carbon dioxide gas is produced
add a piece of magnesium metal to the acid	effervescnece observed, and hydrogen gas is produced
pH meter	pH recorded is below 7
a few drops of acid is added to damp blue and red litmus paper	the blue litmus paper turns red while the red litmus paper remains red

Reactions with acids

-Acid + base = salt + water
-Acid + alkali = salt + water
-Acid + carbonate = salt + carbon dioxide + water
-Acid + metal = salt + hydrogen gas

What are alkalis and bases?

-Bases are chemical substances which include all metal oxides and hydroxides
-Alkalis are substances which dissociate/ionise in water to from OH- or hydroxide ions
-Alkalis are basically soluble bases
-Alkalis are soapy and corrosive when concentrated

-They turn red litmus paper blue
-They have a bitter taste
-They have a pH above 7
-They turn green universal indicator purple if strong and blue if weak
-The stuff on strength and concentration can be applied to alkalis too

Test for alkalis

Test	Result
2-3 drops of universal indicator	aniversal indicator turns from green to purple or blue depending on strength
pH meter	pH recorded is above 7
a few drops of alkali is added to damp blue and red litmus paper	the red litmus paper turns blue while the blue litmus paper remains blue

Reactions with alkalis

-Alkali + acid = salt + water
-Alkali + ammonium salt = salt + ammonia gas + water
-Neutralisation reaction ionic equation: $H^+ (aq) + OH^- (aq) = H_2O (l)$

pH

-pH measures the concentration of H^+ ions in a solution
-Formula for pH: $-\log[H^+]$, where log is log base 10, and H^+ is the concentration of hydrogen ions in mol/dm^3

Types of oxides

Type of oxide	Contains	Reacts with	Examples
acidic (dissolves in water to form acid)	oxides of non metals	base	SO_2, CO_2, NO_2
basic (some dissolve in water to form alkali)	oxides of metals (mostly group 1 and 2)	acid	Na_2O, MgO, CaO
amphoteric	oxides of metals (near the line separating metals and non metals on the periodic table)	both acids and bases	PbO, ZnO, Al_2O_3
neutral (insoluble in	oxides of non metals	nothing	CO, NO, H_2O

water)			

-Amphoteric oxides will act as an acid or an alkali, so if you put it in a solution of both acids and alkalis, they will react with both

Chapter 9: Salts

What are salts?

-A salt is a chemical compound formed by the replacement of one or more hydrogen ions of acids by a metallic ion
-They usually contain at least one metallic cation and one non metallic anion

Acid vs normal salts

-Acid salts contain hydrogen ions
-Examples include sodium bicarbonate ($NaHCO_3$) and sodium bisulfate ($NaHSO_4$)
-Normal salts do not contain hydrogen ions
-Examples include sodium carbonate (Na_2CO_3) and sodium sulfate (Na_2SO_4)
Anhydrous vs hydrated salts

Anhydrous salts	Hydrated salts
do not contain water	contain water
usually a powder	usually a crystal
$CuSO_4$ anhydrous copper (II) sulfate, which is a white powder	$CuSO_4 . 5H_2O$ copper (II) sulfate pentahydrate, which is a blue crystal

-The water in hydrated salts are chemically combined with the salt, and are different from salts dissolved in water

Solubility of salts

Type of salt	Soluble salts	Insoluble salts
nitrates	all nitrates	nil
ammonium	all salts	nil
group 1 (sodium, potassium etc)	all salts	nil
chlorides, bromides, iodides (group 17 halogens)	all except lead and silver	lead and silver

sulfates	all except lead, barium and calcium	lead, barium and calcium (sparingly soluble)
carbonates	group 1, ammonium	all except group 1 and ammonium
hydroxides	group 1, ammonium (calcium, barium and strontium are sparingly soluble)	all except group 1 and ammonium

Preperation of salts

-Salts can be prepared through three different methods, precipitation, titration and reaction of acid with insoluble substances

Precipitation

-Precipitation is used to prepare insoluble salts
-A fixed volume of two soluble salt solutions, where one contains the cation of the salt and one contains the anion of the salt is added and allowed to react
-The mixture is then filtered to obtain the insoluble salt
-The salt is then washed with distilled water and dryed between filter paper

Titration

-Titration is used to prepare group 1 and ammonium salts
-Using a pipette, 25 cm^3 of a known concentration of aqueous alkali is added into a conical flask
-A few drops of indicator is then added to the conical flask
-Dilute acid is added from a burette to the alkali solution, swirling the flask continuously
-Once the indicator changes colour, stop adding acid as the alkali has been neutralised
-Record the volume of acid used
-Repeat the previous steps with the recorded value of acid but without adding indicator
-Heat to evaporate the solution until a third of it is left
-Cool the hot saturated solution and allow crystals to form
-Filter to obtain the crystals, and wash with cold distilled water and dry between filter paper

Reaction of acids with insoluble substances

-When preparing soluble non group 1 or ammonium salts, excess solid metal carbonate, oxide or hydroxide is added to a fixed volume of acid until the solid can no longer dissolve or no more effervescence occurs if carbonate was used
-Filter the mixture to remove the excess solid
-Heat to evaporate the solution until a third of it is left
-Cool the hot saturated solution and allow crystals to form
-FIlter to obtain the crystals, and wash with cold distilled water and dry between filter paper

Chapter 11: Qualitative Analysis

What is qualitative analysis?

-Qualitative analysis is a process of identification of cations and anions in an unknown salt
-This can be done through various tests, such as tests for gases, cations, anions, solubility, heating,oxidisng agent and reducing agent

Colour of substances

Colour	Possible identity
white solid	ammonium salts and most group 1-3 substances
black solid	copper (II) oxide, iron (II) oxide
grey solid	most solid metals, such as iron, magnesium and zinc
blue solid	most copper (II) salts, such as copper (II) sulfate and copper (II) nitrate
green solid	iron (II) salts, some copper (II) salts such as copper (II) carbonate
brown solid	iron (III) salts, copper metal

How to record a produced gas

-If a gas is liberated in a chemical reaction, the following must be recorded
-Observations (effervescence, colour, odour)
-How the gas was tested and result of the test
-Identity of gas from test

Test for gases

Gas	Colour and odour	Test	Observation	Nature of gas
H2	colourless and odourless	place a lighted splint at the mouth of the test tube	lighted splint extinguished with a pop sound	neutral
O2	colourless and odourless	insert a glowing splint into the test tube	gas relights glowing splint	neutral
CO2	colourless and odourless	bubble gas into limewater	gas forms white precipitate in	acidic

			limewater; if more carbon dioxide is bubbled into limewater, white precipitate dissolves	
Cl2	green and pungent	place moist litmus papers at the mouth of the test tube	gas turns moist blue litmus paper red and bleached/white	acidic and bleaching
SO2	colourless and pungent	place filter paper dipped in purple potassium manganate (VII) at the mouth of the test tube	gas turns purple potassium manganate (VII) colourless	acidic and reducing agent
NH3	colourless and oungent	place moist litmus papers at the mouth of the test tube	gas turns moist red litmus paper blue	alkaline

How can cations be identified?

-A cation in an aqueous salt can be identified using sodium hydroxide and aqueous ammonia
-These react with the salt to from the hydroxide of the cation, which is usually a precipitate
-The following must be recorded when testing for cations
-The colour of precipitate formed; if any
-Whether the precipitate is soluble/insolublein excess reagent (about twice the original volume of the sample)

Test for cations

Cation	NaOH	Excess NaOH	NH3 (aq)	Excess NH3 (aq)	Respective hydroxide/precipitate
group 1	no precipitate	nil	no precipitate	nil	group 1hydroxide
NH4+	no precipitate, ammonia gas produced on warming	nil	no precipitate	nil	NH4OH
Cu2+	blue precipitate	insoluble in excess	blue precipitate	soluble in excess to form	Cu(OH)2

				a dark blue solution	
Fe^{2+}	dirty green precipitate, turns brown when left standing	insoluble in excess	dirty green precipitate, turns brown when left standing	insoluble in excess	$Fe(OH)_2$
Fe^{3+}	reddish brown precipitate, forms when $Fe(OH)_2$ is left standing	insoluble in excess	reddish brown precipitate, forms when $Fe(OH)_2$ is left standing	insoluble in excess	$Fe(OH)_3$
Ca^{2+}	white precipitate	insoluble in excess	no precipitate	nil	$Ca(OH)_2$
Zn^{2+}	white precipitate	soluble in excess to from a colourless solution	white precipitate	soluble in excess to from a colourless solution	$Zn(OH)_2$
Al^{3+}	white precipitate	soluble in excess to from a colourless solution	white precipitate	insoluble in excess	$Al(OH)_3$
Pb^{2+}	white precipitate	soluble in excess to from a colourless solution	white precipitate	insoluble in excess	$Pb(OH)_2$

-To distinguish aluminum salts from lead (II) salts, potassium iodide is added to both samples
-A yellow precipitate is formed with the lead (II) salt but no visible reaction will be seen with the aluminum salt
-In general, the ionic equation of the cation test is: A^{n+} (aq) + OH^- (aq) = $A(OH)_n$ (s)

Test for anions

Anion	Test	Observation	Chemical/ionic equation
CO_3^{2-}	add dilute acid and test the gas evolved	effervescence is observed, gas evolved forms white precipitate in limewater, gas is CO_2	1. CO_3^{2-} (aq) + $2H^+$ (aq) = CO_2 (g) + H_2O (l) 2. CO_2 (g) + $Ca(OH)_2$ (aq) = $CaCO_3$ (s) + H_2O (l)
Cl^-	add dilute HNO_3,	a white precipitate of AgCl	Ag^+ (aq) + Cl^- (aq) = AgCl (s)

	followed by $AgNO_3$ or add dilute HNO_3, followed by $Pb(NO_3)_2$	forms or a white precipitate of $PbCl_2$ forms	or $Pb^{2+} (aq) + 2Cl^- (aq) = PbCl_2 (s)$
I^-	add dilute HNO_3, followed by $AgNO_3$ or add dilute HNO_3, followed by $Pb(NO_3)_2$	a yellow precipitate of AgI forms or a yellow precipitate of PbI_2 forms	$Ag^+ (aq) + I^- (aq) = AgI (s)$ or $Pb^{2+} (aq) + 2I^- (aq) = PbI_2 (s)$
NO_3^-	add alkali follwed by a piece of aluminum; warm and test for gas evolved	gas evolved turns moist red litmus paper blue, gas is NH_3	1. $NO_3^- (aq) = NH_4^+ (aq)$ 2. $NH_4^+ (aq) + OH^- (aq) = NH_3 (g) + H_2O (l)$
SO_4^{2-}	add dilute HNO_3, followed by $Ba(NO_3)_2$ or add dilute HCl, followed by $BaCl_2$	a white precipitate of $BaSO_4$ forms	$Ba^{2+} (aq) + SO_4^{2-} (aq) = BaSO_4 (s)$

-Acid is added to the test for chlorides, iodides and sulfates to test for carbonates, as they can give false results