Chemistry RP Year 3 Notes

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Topic 1 – Elements, Compounds, Mixtures and the Particulate Nature of Matter

Substances can be categorized into 3 main types:

1. Elements
2. Compounds
3. Mixtures

Characteristics of Elements

- Simplest Substance
- Cannot be broken down further by chemical methods
- An atom of an element is a particle representing an element that is not chemically bonded with other atoms
- A molecule of an element is a particle representing an element that is chemically bonded with other atoms of the same kind
- Elements are considered pure
- 1 type of atom
- 1 substance

Characteristics of Compounds

- Molecule of a compound is a particle representing a compound with 2 or more atoms chemically bonded together
- Atoms must be of 2 or more different kinds (elements)
- 1 substance

Characteristics of Mixtures

- Can consist of 2 or more elements not chemically bonded together – E.g. O_2 and H_2 mixture
- Can consist of 1 or more elements not chemically bonded with 1 or more compounds – E.g. Air
- Can consist of 2 or more compounds not chemically bonded together – E.g. Salt Water ($NaCl$ + H_2O

Similarities and Differences between Elements and Compounds

Similarities	Differences
Both are made up of particles (atoms) which are the building blocks	Elements are only made up of one type of atom while compounds are made up of more than one type of atom
Both are pure substances	An element may be metallic or non-metallic while a compound can only be non-metallic, while either being made of metallic and non-metallic elements or made entirely of non-metallic elements
Both have fixed melting and boiling points	An element cannot be chemically split into simpler substances while a compound can be decomposed into its elements or simpler compounds

Similarities and Differences between Compounds and Mixtures

Similarities	Differences
Both are made of more than one type of atom	A mixture can be separated into its components by physical methods (e.g. filtration, distillation) while a compound can only be separated into its elements via chemical methods
	A mixture does not have a fixed boiling or melting point while a compound does
	No chemical changes take place when a mixture forms while a chemical change takes place when a compound forms
	The proportion of the constituents in a mixture can vary, while elements are always combined in a fixed proportion in a compound
	The properties of a mixture are the same as those of its constituents, while the properties of a compound differ from those from its constituents

Simple definitions of the 3 types of substances:

- Element – A substance that is made up of only one type of atom. It cannot be chemically split into simpler substances.
- Compound – A substance that contains two or more elements chemically joined together.
- Mixture – A substance that contains two or more substances physically together but not reacted with one another chemically.

The Particulate Nature of Matter

The Kinetic Theory states that:

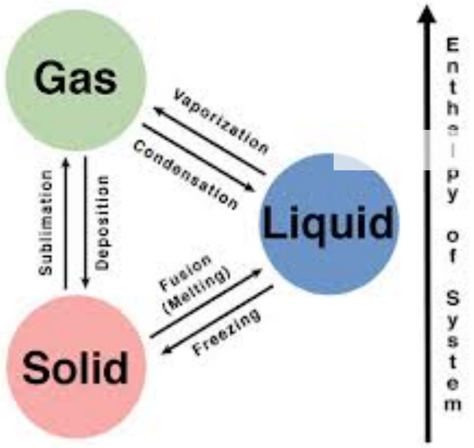
1. All matter is made of atoms, the smallest unique particle of each element. A particle of a gas could be an atom or a group of atoms (molecules).
2. Atoms/molecules have some energy of motion that we feel as temperature. The motion of atoms/molecules can be in the form of vibration and/or linear motion of translation.
3. The pressure of a gas is due to the motion of the molecules of gas striking the object bearing that pressure.
4. There is a very large distance between the molecules of a gas compared to the size of the molecule such that the size of the molecule can be considered negligible.

States of Matter

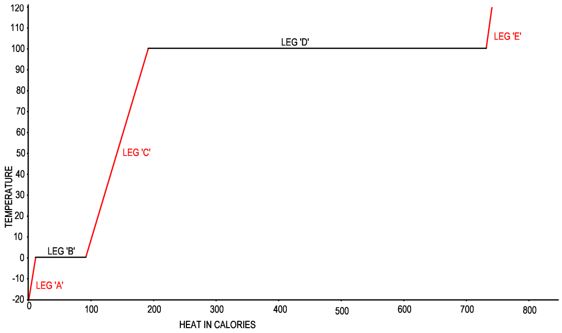
	Solid	**Liquid**	**Gas**
Energy Content	Low	Moderate	High
Arrangement of Particles	Close together in an orderly arrangement	Close together but not in an orderly arrangement	Far apart and not in an orderly arrangement
Motion	Vibration	Vibration and Translation	Vibration and Translation
Shape and Volume	Has a fixed shape and volume	Has no fixed shape (takes the shape of the container) but has a fixed volume	Has no fixed shape and no fixed volume (fills the volume of the container)
Forces of attraction between particles	Strong	Moderate	Weak or None
Compressible	No	No	Yes

*Note: Use ADAMS (Arrangement of particles, Distance between particles, forces of Attraction between particles, Movement of particles, and Speed of movement of particles) when describing the particulate nature of solids, liquids or gases.

Phase Transition



Below is the Heat Curve illustrating corresponding changes in energy and temperature of water as it undergoes a phase transition between liquid and gas.



Therefore, at 100°C, water will undergo a phase transition into the gas phase. At this point, energy introduced to the liquid will only go to sending water molecules into the gas state.

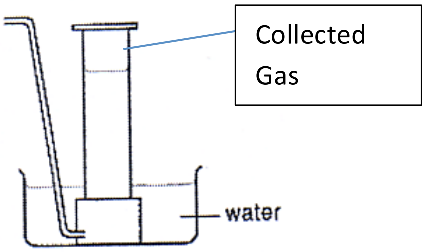
Topic 2 – Experimental Design and Purification Techniques

Production, Collection and Testing of a Gas

The usual apparatus for measuring the volume of gas is the gas syringe.

Collection of a gas depends on its solubility in water and its density compared to that of air.

Downward water displacement can be used if the gas is insoluble (E.g. Hydrogen, Methane). Note that the gas is pure when collected.



Downward delivery is used if the gas is denser than air (E.g. CO_2, Cl_2, Br_2).

Upward delivery is used if the gas is less dense than air. (E.g. He)

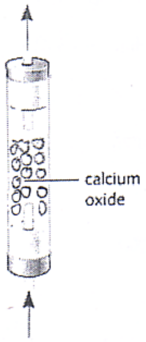
Drying of Gases (Removal of moisture)

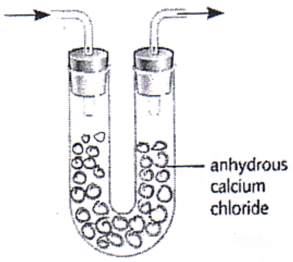
Gases are usually contaminated with water vapour. Hence, the mixture is passed through a drying agent to dry the gas (remove moisture).

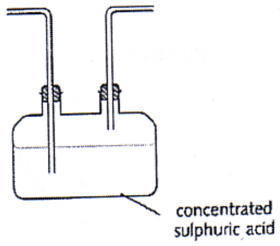
There are 3 types of drying agents:

1. Basic/Alkaline drying agents (E.g. CaO) – Cannot dry acidic gases, such as CO_2
2. Acidic drying agents (E.g. H_2SO_4) – Cannot dry basic/alkaline gases, such as NH_3 (ammonia)
3. Neutral drying agents (E.g. $CaCl_2$) – Can dry neutral, acidic, and basic gases

There are 3 setups that can be used for drying:







The first setup can ONLY be used for solid drying agents.

The second can be used for both solid and liquid drying agents.

The last one can ONLY be used for liquid drying agents.

After these processes, the gas can be tested.

Purification Techniques

1. Dissolving, Filtering and Evaporating – Used for a mixture of soluble and insoluble impurities in the substance. Steps are to extract the solute ONLY.
   E.g.: A mixture of sand and NaCl only.
2. Crystallization –

Take note that water molecules must be present because they are what give the crystal its shape. As such, excessive heating would result in crystals not forming. (*NOTE: General rule of thumb is to heat the solution up on an evaporating dish until only 1/3 of the solution remains because this is when the solution is saturated, and then allow the solution to cool)
E.g.: Crystallization of Cu(II)SO_4 (Copper (II) Sulfate) from its solution

3. Sublimation – Some solid substances become gaseous when heated under room conditions, without going through the liquid state. Examples include NH_4Cl (ammonium chloride), I_2 (iodine), and dry ice (CO_2). Heat the mixture in an evaporating dish and invert a filter funnel

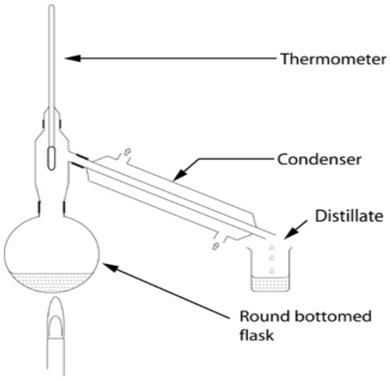
over the dish. The sublimed solid will solidify as a deposit on the cooler inner surface of the funnel.

E.g.: Mixture of NaCl and NH_4Cl only.

4. Separating funnel – When two immiscible liquids have to be separated.

   E.g.: Oil and water

5. Simple distillation – Used to obtain solvent from a solution



IMPORTANT NOTES

1. Broken porcelain is used as boiling chips to ensure the smooth boiling by allowing formation of smaller air bubbles.

2. The thermometer is placed at the exit of the flask to measure the temperature of the pure vapour exiting the flask, and hence determines the boiling point of the solvent.

3. The cooling water enters the condenser from the bottom to ensure that the bottom is the coolest part, so any vapour that did not condense earlier would condense here, and so no vapour can escape.

E.g.: Get water from salt solution.

6. Fractional Distillation – Used when separation of a mixture of miscible liquids with different boiling points is required. Note that the liquid with the lower boiling point will be collected first, and during that evaporation process, the reading on the thermometer will remain constant. After completion, the reading will begin to rise again.

IMPORTANT NOTES:

1. Do not use a Bunsen burner if flammable liquids must be separated (naked flames must not be used to heat flammable liquids); instead use an electric heater.

2. The tiny glass beads in the fractionating column are there to increase the surface areas for the condensation of vapour, making the separation of the 2 vapours more efficient.

3. The reading will first record the boiling point of the liquid which has a lower boiling point, and after all the vapour of that liquid has entered the condenser the thermometer will then record the boiling point of the other liquid.

E.g.: Separate a mixture of water and ethanol (100°C vs. 78°C)

7. Reverse Osmosis – High pressure is used to force impure water through a membrane which has millions of tiny pores. Only water molecules can pass through the pores, leaving solute particles behind.

Tests for Purity of a Substance

The presence of impurities in a substance decreases the melting point of a substance and causes it to melt over a range of temperatures. It also increases the boiling point and causes it to boil over a range of temperatures.

You can also test for the purity through the use of the Rf value. (Rf = Dist. Travelled by Solute/Dist. Travelled by Solvent)

Identification of Gases

Chlorine Test – Look out for a pale yellow/green gas that has evolved and test it with a piece of moist blue litmus paper. The paper should turn red and then should become bleached.

Hydrogen Test – Effervescence should be observed. When inserting a lighted wooden splint, the gas extinguishes the lighted splint with a pop.

Oxygen Test – Effervescence should be observed. When inserting a glowing splint, the gas rekindles the glowing splint.

Sulfur Dioxide Test – Colourless and choking gas should have evolved.

Carbon Dioxide Test – Effervescence observed. A colourless gas is evolved. A white precipitate is formed in limewater.

Ammonia Test – Colourless and pungent gas evolved and when tested with moist red litmus paper, it turns blue. Effervescence should also be observed.

SPA – Look through W/S 1 & 2

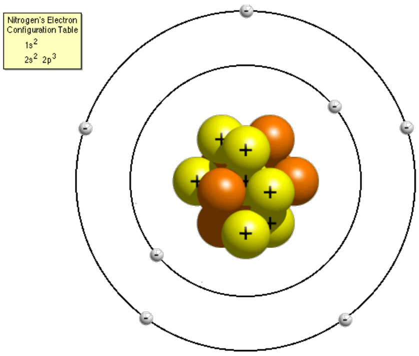
NOTE: Ensure you can write steps for experiments, safety precautions, and provide the theoretical basis of why a certain method is used to solve the question.

Topic 3 – Atomic Structure

Models of the Atom

The Bohr Model

Electrons can only exist and travel in special energy levels termed ‘shells’. These shells are at fixed distances from the nucleus and electrons within each shell have a specific energy. These electrons can only gain or lose energy from jumping from one shell to the next (no continuous energy gaining/losing) in packets of energy called ‘quanta’.



A fixed number of electrons can fit into each shell. This can be found using the formula $N = 2n^2$, where n is the shell number. For example, in the 2nd shell, the number of electrons that can fit is $2 \times 2^2 = 8$.

Problems

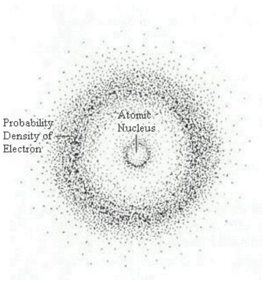
However, there are inconsistencies with the model. According to the model, for example, potassium (K) is meant to have an electron configuration of 2.8.9. However, potassium instead has an electron configuration of 2.8.8.1. Many other atoms deviate from this rule, and so other models began to supersede the Bohr model.

Electron Cloud Model

Electrons have particle-like and wave-like properties. Hence, the position of the electron cannot be exactly known, and so the wave-mechanical models of the atom were proposed.

The electron cloud describes electron locations in terms of probable regions, with fuzzy boundaries – hence the term 'cloud'.

The following diagram uses the electron cloud model to represent an atom.



The positive nucleus is surrounded by a cloud of electrons 'floating' in their energy levels. Lower energy electrons are in levels nearer to the nucleus while higher energy electrons are in levels further away from the nucleus.

Isotopes

Definitions

Isotopes: Atoms of the same elements but of different masses (different number of neutrons). This is because they have a different nucleon/mass number.

Characteristics

Isotopes have the same number of electrons. Hence, they will have the same chemical properties (Chemical properties depend upon the transfer and redistribution of electrons). However, as isotopes have different mass numbers, they will have different masses, and hence will have different physical properties. For example, deuterium, which is found in hard water, is an isotope of hydrogen, and is represented as ${}^{2}_{1}H$. As it is heavier than a hydrogen atom, hard water, which is made of 2 deuterium atoms and one oxygen atom, has a melting point of 3.8°C, while normal water has a melting point of 0°C.

Relative Atomic Mass

The relative atomic mass of any element is the average relative mass of one atom, taking into account the different isotopes and their RELATIVE PROPORTION.

The reference mass for the relative atomic mass is $\frac{1}{12}$ of a carbon-12 atom or the mass of a hydrogen atom. Hence, essentially the reference mass is the mass of one nucleon (protons and neutrons have the same mass).

They can be calculated using the formula:

$$A_r = I_1 \times x\% + I_2 \times y\%,$$

Where I_1 is the mass number of the first isotopes, x is the abundance by percentage of the isotope, I_2 is the mass number of the second isotope, and y is the abundance by percentage of the second isotope.

*Don't remember it using the formula – It is plain logic.

Radioisotopes

There are two types of isotopes – radioactive isotopes and non-radioactive isotopes. Radioactive isotopes are called radioisotopes. Radioisotopes are rarely found naturally, and are found in small amounts. Most radioisotopes are made artificially.

Radioactive isotopes have important uses; cobalt-60 is used in hospitals for radiotherapy for cancer patients, as the radiation given off kills cancer cells. Cobalt-60 is also used for sterilization of surgical instruments.

Uranium-235 is used as nuclear fuel in nuclear power stations because it undergoes nuclear fission.

Isotopes and the Mass Spectrometer

The mass spectrometer allowed scientists to discover isotopes.

The mass spectrometer consists of 3 basic components; the ioniser, the mass analyzer, and the detector.

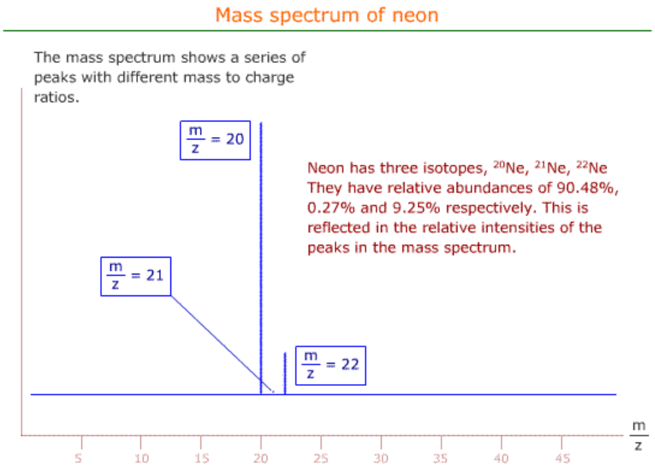
Steps

A gaseous sample of an element is first bombarded with electrons or laser in order to create cations of the element.

In the mass analyzer, ions of the same charge are separated according to their mass to charge ratio (m/z) by using a magnetic or electric field.

The data obtained using the detector is then presented as a mass spectrum in which m/z is plotted against the intensity of the respective ion signals.

The following is a mass spectrum of neon:



Mass spectrum of neon

The mass spectrum shows a series of peaks with different mass to charge ratios.

Neon has three isotopes, ^{20}Ne, ^{21}Ne, ^{22}Ne They have relative abundances of 90.48%, 0.27% and 9.25% respectively. This is reflected in the relative intensities of the peaks in the mass spectrum.

Mass spectrometry is applied in many fields, including medicine, forensic science, the food industry and security. It is also used to determine radiocarbon dating in archaeology.

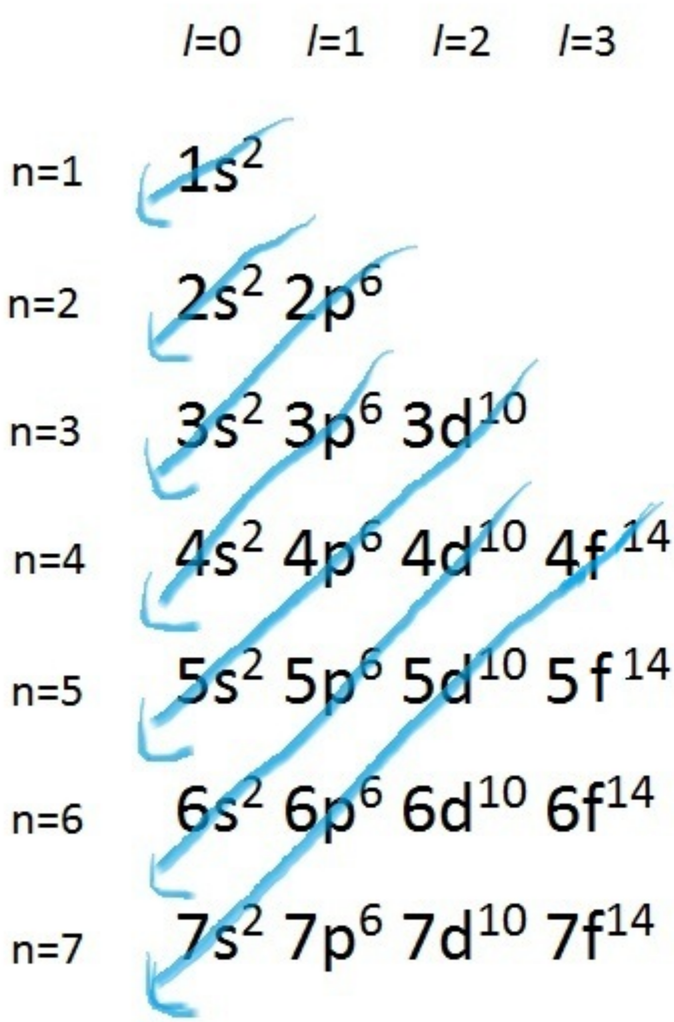
Subshells and Orbitals

Shells are comprised of different subshells. There are 4 types of subshells: the s, p, d, and f subshells. Each type of subshells contains a different number of orbitals.

Type of subshell	Number of orbitals
s	1
p	3
d	5
f	7

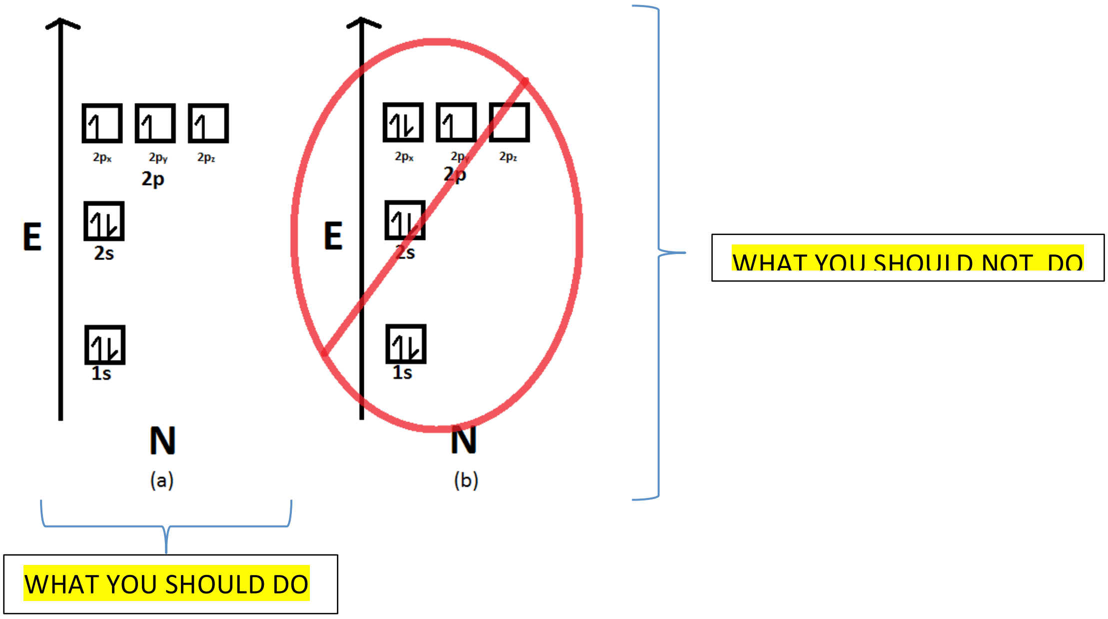
*Note: Each orbital can be filled only by two electrons (with opposite spins).

The subshells are filled according to energy level in the following order:



*Note: When filling up orbitals within either a p, d, or f subshell, ensure that every orbital has one electron before adding the second electron to each subshell. For example, in a nitrogen atom, the 1s subshell is filled, followed by the 2s subshell. In the 2p subshell, the $2p^x$ orbital is first filled with one electron, then the $2p^y$ orbital, and $2p^z$ orbital. DO NOT fill the $2p^x$ orbital entirely, and then place the next electron into the $2p^y$ orbital.

You might need to represent it using a box diagram.



Electronic Configuration

The electronic configuration, by convention, must be written in the order of the shells. E.g.: Electronic configuration of selenium = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^4$. However, the order in which the subshells are filled remains the same; this means certain elements that have electrons in the 4s subshells do not have electrons in the 3d subshell (Example would be calcium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$).

Hence these are the steps to follow for writing the electronic configuration:

Step 1: Write down electronic configuration according to energy levels of subshells (E.g. selenium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ – Difference is highlighted)

Step 2: Rearrange it according to shell number.

Atomic Structure and Ionisation Energy

The electronic structure of an atom refers to the way its electrons are arranged in the atom. The study of the ease with which atoms lose electrons allows scientists to find useful information about the arrangement of these electrons. The energy required to remove one electron from a gaseous atom is known as its **ionisation energy (IE)**.

Example:

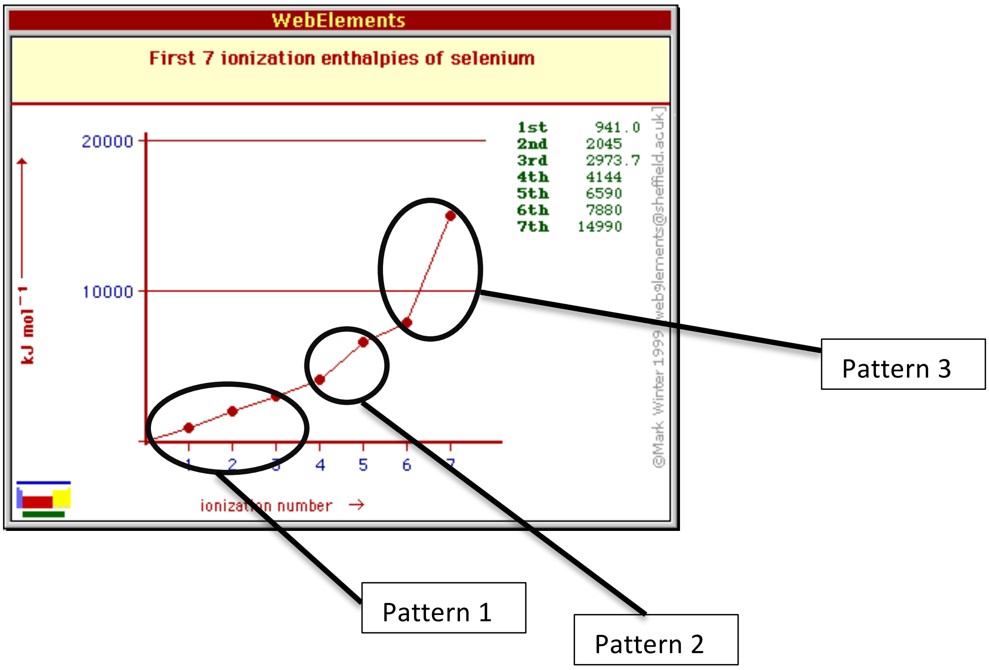
The first ionisation energy of sodium (Na) is the energy required to remove the first electron from a neutral sodium atom, to form a sodium ion (Na^+). The second ionisation energy is the energy required to remove the second electron from the Na^+ ion. Additional supplies of energy will result in the removal of successive electrons – a succession of ionisations is hence possible.

Patterns to note:

1. The ionisation energies increase for each successive loss of electrons

2. There is a significant jump in the ionisation energy whenever an electron is removed from a shell closer to the nucleus than the previous electron removed
3. There is a modest jump in the ionisation energy whenever an electron is removed from a subshell closer to the nucleus than the previous electron removed

Graph of ionisation energies of selenium:



Topic 4 – Chemical Bonding & Structure

Octet Structure

The octet structure refers to the understanding that most atoms, when achieving noble gas configuration to attain stability, tend to have 8 electrons in their outermost shell (s and p subshells filled).

Most elements tend to obey this rule, although a few do not. The first few elements in a periodic table tend not to obey this rule as they instead achieve a duplet structure, which is that of helium. “Octet-expanders”, which are elements that are in Period 3 and beyond, can accommodate more than 8 electrons in their valence shell, though in the end, the noble gas configuration only contains 8 electrons in their valence shell.

*Note: Do not use “fully-filled” shell anymore; instead employ the phrase “achieve(s) an octet structure”.

Ionic Bonding

Definition: The electrostatic forces of attractions between ions of positive and negative charges.

Characteristics of Ionic bonding

1. Occurs between metallic elements and non-metallic elements
2. Metal atoms (Group I, II, III) tend to lose electrons to form cations (positively charged ions)
3. Non-metal atoms (Group V, VI, VII) tend to gain electrons to form anions (negatively charged ions)

Example

Sodium (Na) reacts with chlorine (Cl) to form sodium chloride.

Sodium

Net charge of sodium atom = $(+11)+(-11)=0$

When the sodium atom loses an electron to achieve the octet structure of 2.8, it becomes a cation.

Net charge of sodium ion = $(+11)+\big(-(11-1)\big)=(+11)+(-10)=1+$

Chlorine

Net charge of chlorine atom = $(+17)+(-17)=0$

When the chlorine atom gains an electron to achieve the octet structure of 2.8.8, it becomes an anion.

Net charge of chloride ion = $(+17)+\big(-(17-1)\big)=(+17)+(-16)=1-$

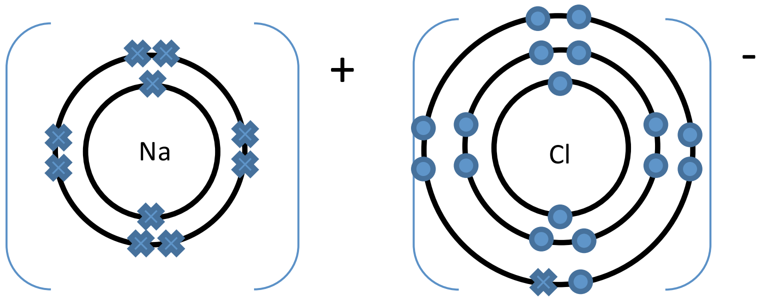
These oppositely charge ions attract each other, and hence are held together by electrostatic forces of attraction.

*Note: One cation does not bond only to one anion. The electrostatic forces of attractions are present between any cation and anion. Hence, one cation is attracted to multiple anions and vice-versa. This means that NO molecules are present in an ionic substance.

Dot-and-Cross Diagrams

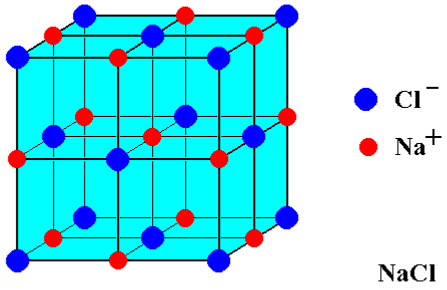
A dot-and-cross diagram is used to represent the electron arrangements within ions and to show the ratio of cations to anions. The diagram, however, does NOT imply that an ionic bond is only formed between the ions shown.

The following is a dot-and-cross diagram of sodium chloride (NaCl), showing the sodium cation and chloride anion.



A more complete idea is to view ionic bonding in terms of its function in ionic structures. In ionic solids, the ions are arranged into a giant crystal lattice, with many ionic bonds between the oppositely charged ions.

The diagram below shows the ionic lattice structure of NaCl.



Notice how one Na^+ cation is bonded to 6 Cl^- anions, and vice versa.

Covalent Bonding

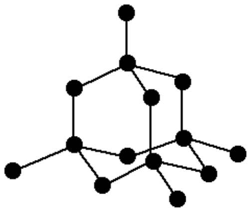
Definition: Sharing of electrons in valence shell to form a molecule.

Characteristics of Covalent Bonding

1. Occurs only between non-metals
2. Can occur between atoms of the same element or atoms of different elements
3. Can have single bonds, double bonds or triple bonds

Types of Covalent Bonding

Macromolecular Structures (M.M.S)

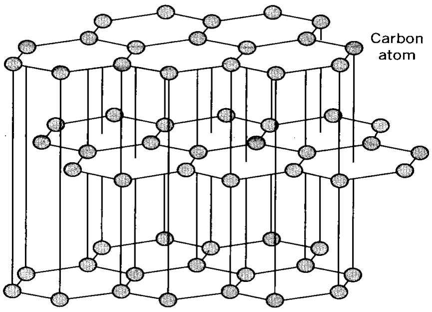


The above diagram is a model of diamond, which is made up of carbon atoms covalently bonded to one another. The covalent bonding extends throughout the tetrahedral network structure.

Hence, diamond has high melting and boiling points, and is very hard. It also is a non-conductor of electricity in any state.

However, graphite is a macromolecular structure that is slightly different.

The following is a model of graphite.



Graphite is comprised of multiple layers of covalently bonded carbon atoms, with one carbon atom bonded to another 3. There are delocalized electrons present in each layer, from each carbon atom. The covalent bonds extend throughout the network structure of each layer. However, the layers are held together by weak intermolecular forces (Van der Waals' forces).

Hence, graphite has high melting and boiling points. It, however, unlike other macromolecular covalent substances, conducts electricity in the direction parallel to the layers, due to delocalized electrons being present in each layer.

Simple Molecular Structures (S.M.S)

Simple molecular structures refer to the molecules of covalent substances. In simple molecular structures, the covalent bonding exists only within the molecule, and there are weak intermolecular forces of attraction between molecules (Van der Waals' forces). Hence, most S.M.S substances exist in either liquid or gaseous states at room temperature. Their melting and boiling points are also relatively low. They do not conduct electricity in any state (with the exception of HCl, which conducts electricity in the aqueous state as it disassociates to form H^+ and Cl^- ions).

Dot-and-cross diagrams

Dot-and-cross diagrams are used to show the bonding between atoms in a molecule.

The following is a dot-and-cross diagram of carbon dioxide.

Notice how only the valence electrons are shown, and how shared electrons are between the atoms that are sharing those electrons.

*Note: Covalent bonding is not a weak bond. It is a very strong bond, just like ionic bonds. However, the melting and boiling points of S.M.S are low because during melting and boiling, only the weak intermolecular forces between molecules are overcome, and not the strong covalent bonds within the molecule itself.

Dative/Coordinate Bonding

Definition: A covalent bond in which only one atom provides the electrons for sharing. That atom is known as the donor, while the other that does not share is termed the acceptor.

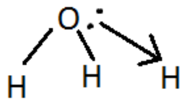
Conditions: The donor must have at least one unshared pair of electrons, while the acceptor must have at least one vacant orbital within the shell.

For example, in water, the oxygen atom within the H_2O molecule has 2 unshared pairs of electrons. This can be shared with an atom which needs them to complete its valence shell. A proton, H^+, has one empty 1s orbital. Hence, the H^+ accepts a pair of electrons from oxygen, thus achieving a full first shell.

Equation: $H_2O + H^+ \rightarrow H_3O^+$

The polyatomic ion formed on the right side of the equation is known as a hydronium ion, and has a charge of 1+.

*Note: When drawing the structural formula of any molecule with a dative bond, ensure that instead of a line, an arrow is drawn from the donor to the acceptor. There should be 2 dots, representing a pair of electrons, behind the arrow.)



Electronegativity

Definition: A measure of the tendency of an atom to attract a bonding/shared pair of electrons.

Units: Pauling scale

When 2 atoms of equal electronegativity bond together, the electrons are more or less equidistant from both atoms. This occurs primarily in diatomic molecules, such as H_2 and Cl_2.

When 2 atoms of different electronegativity bond together, the electrons are more attracted to the atom with a greater electronegativity. Hence, the electrons will be found, on average, closer towards the atom with a greater electronegativity. As such, the atom which is further away from the shared pair will have a slightly positive charge, and the one which is closer will be slightly negative. This is represented using the $\delta +$ and $\delta -$ symbols (read as “delta-plus” and “delta-minus”).

When the 2 atoms have a vast difference in electronegativity, the electrons are so attracted to the atom with the greater electronegativity that the atom has complete control over the electrons and the other atom does not have any control. As such, ions form.

We can predict the “ionic”/ “covalent” nature of a bond using the electronegativity difference (ΔEN) between the two bonding atoms.

$0 \leq \Delta EN < 0.5$	Non-polar covalent
$0.5 \leq \Delta EN < 2.0$	Polar covalent
$2.0 \leq \Delta EN \leq 4.0$	Ionic

Polarizability

An anion that readily undergoes a large distortion of its electron cloud is said to be highly polarizable, while a cation that can cause such large distortions is said to have a high polarizing power.

Larger anions are generally more polarizable, which is why I^- (54 electrons) is more polarizable than Br^- (36 electrons).

Smaller cations that are highly charged are generally higher in polarizing power. Hence, the bond formed between Cl^- and Al^{3+} in $AlCl_3$ is covalent, although the elements involved are non-metallic and metallic respectively. This is because the strong polarizing power of Al^{3+} causes a pair of electrons from Cl^- to move so close to the Al^{3+} ion that the bond is mainly covalent in nature.

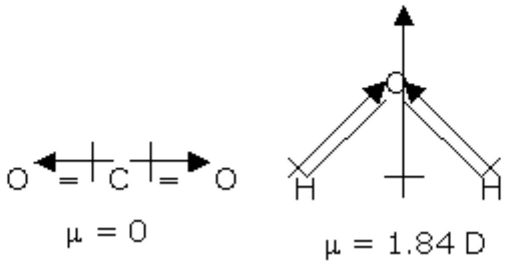
Dipole

Definition: A dipole is a separation of opposite charges. Hence, in a polar bond that connects one slightly positive atom to another slightly negative atom, a dipole in present.

Quantity: The size of a dipole is a dipole moment, which is measured in units called Debye (D) or coulomb meter (C.m).

Dipoles are directional, from the electropositive atom to the electronegative atom. As such, they can cancel each other out. In CO_2, the dipole moment is in opposite directions. Hence, they cancel each other out, and so CO_2 is a non-polar molecule.

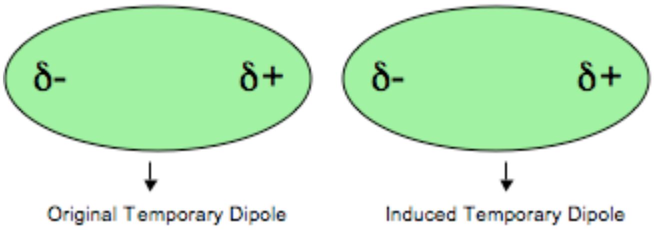
However, in H_2O, the dipole moments do not lie opposite to each other but at an angle. Hence, they do not cancel, and so H_2O is a polar molecule.



*Note: In general, the dipole moments in a symmetrical molecule tend to cancel each other out, resulting in a non-polar molecule, while an asymmetrical molecule with dipoles tends to be a polar molecule. (A dipole is represented with an arrow that has a '+' tail that points towards the negative partial charge.)

Instantaneous dipole - Induced dipole (id-id) interaction/Van der Waals' forces

As electrons in a molecule are in continual motion, at any particular moment, the electron cloud of the molecule will not be perfectly symmetrical, and there will be more negative charge on one side than on the other. This causes a temporary or instantaneous dipole. This induces dipoles in neighbouring molecules.



*Note: id-id interactions are short-lived as the electrons are continually in motion and so the dipoles may vanish and reform. Hence, the overall attraction tends to be weak.

Factors affecting strength of id-id interactions

1. The number of electrons in the molecule – A larger number of electrons means a larger electron cloud, and so the greater the id-id interaction
2. The greater the surface area of the molecules the greater the id-id interaction

Stronger id-id interactions lead to a higher boiling and melting point.

Permanent dipole-Permanent dipole (pd-pd) interactions

Polar molecules have permanent dipoles. Hence, they tend to align in a manner such that the slightly positive end of one molecule is near the slightly negative end of another molecule.

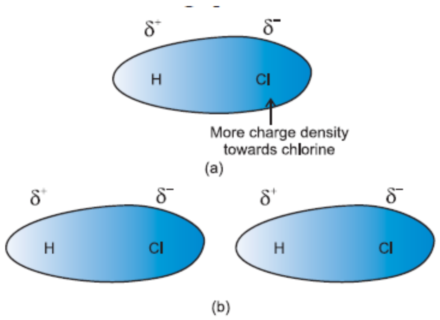


Fig. 5.2 (a) Distribution of electron cloud in HCl – a polar molecule, (b) Dipole-dipole interaction between two HCl molecules

The electrostatic forces of attraction between the slightly positive end and the slightly negative end of different molecules results in the pd-pd interaction. As the dipoles are permanent, pd-pd interactions are generally stronger than id-id interactions.

Polar molecules hence experience both pd-pd and id-id interactions, making their boiling and melting points higher than those of non-polar molecules.

Hydrogen bonding

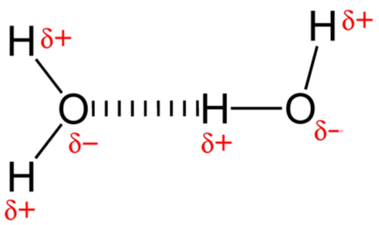
Hydrogen bonding, or H-bonding, is a specific type of pd-pd interaction.

For it to occur between two molecules, one must have a hydrogen atom bond to N, O or F, while the other must have N, O, or F.

The bond is the electrostatic force of attraction between the $H^{\delta+}$ and the lone pair of electrons on N, O or F. This is because the hydrogen atom has no inner core electrons. When bonded to a highly electronegative atom X (N, O or F), the shared electrons are drawn towards X, causing the proton in the hydrogen nucleus to be partially exposed. This nearly bare proton interacts strongly with the lone pair of an electronegative atom Y (also N, O or F) in a nearby molecule, resulting in H-bonding.

Y must be N, O or F as Y must be both small and highly electronegative, and only these 3 atoms satisfy the criteria.

An example would be between H_2O molecules.



These molecules also have pd-pd interactions and id-id interactions, as they have permanent dipoles and instantaneous dipoles. H-bonding is the strongest of all the 3 intermolecular forces.

Metallic Bonding

Definition: The electrostatic forces of attraction between delocalized electrons and positively charged metal ions in a metal.

Example

In a piece of sodium, each atom's valence electron can break free from the sodium atom, forming sodium cations, and wander free throughout the metal. This electron is known as a delocalized electron. An array of sodium cations are hence embedded in a sea of freely moving delocalized electrons from the original sodium atoms. The metal is hence held together due to the electrostatic forces of attraction between the seas of delocalized electrons and cations.

Characteristics of metals

Metals are malleable (can be beaten into sheets) and ductile (can be twirled into wires). This is because their atoms are arranged in layers which can easily slide over one another.

Metals can also conduct electricity as the delocalized electrons can move freely throughout the metal.

Topic 5 – Chemical Formulae and Equations

Types of Chemical Formulae

There are 3 types of chemical formula for any compound: Molecular formula, empirical formula and structural formula.

Molecular formula

If a compound is made up of molecules, then its molecular formula shows all the atoms within the molecule.

Example: The molecular formula of water is H_2O, which means that there are 2 hydrogen atoms and 1 oxygen atom in a molecule.

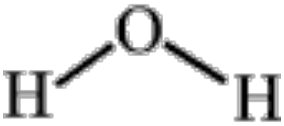
Empirical formula

Empirical formula refers to the simplest formula. The empirical formula of a compound shows the simplest ratio of the elements present. Hence, the empirical formula of glucose, whose molecular formula is $C_6H_{12}O_6$, is CH_2O. For substances not made up of molecules, the commonly used formula, such as in the ionic compound NaCl, is actually the empirical formula.

Structural formula

The structural formula of a molecule shows clearly how the atoms are bonded to one another. Each atom is represented by its alphabetical symbol and bonds are represented by straight line.

Example: Water (H_2O)



Balancing Chemical Equations

Things to note:

1. When no number is written in front of the formula, it is assumed to be the number "1".
2. The number of atoms of each element/type on the left side of the equation must be equal to the number of atoms of each element on the right side
3. Substances of the left side of the equation are called 'reactants' while those on the right are called 'products'.

Formulae of Common Ions

Name of cation	*Formula of ion*	*Name of anion*	*Formula of ion*
Sodium	Na^+	Chloride	Cl^-
Potassium	K^+	Hydroxide	OH^-
Lithium	Li^+	Iodide	I^-
Ammonium	NH_4^+	Nitrate	NO_3^-
Silver	Ag^+	Nitrite	NO_2^-
Hydrogen	H^+	Manganate (VII)	MnO_4^-
Barium	Br^{2+}	Hydrogen Carbonate	HCO_3^-
Calcium	Ca^{2+}	Carbonate	CO_3^{2-}
Copper (II)	Cu^{2+}	Dichromate (VI)	$Cr_2O_7^{2-}$
Iron (II)	Fe^{2+}	Oxide	O^{2-}
Lead (II)	Pb^{2+}	Sulfate	SO_4^{2-}
Magnesium	Mg^{2+}	Sulfite	SO_3^{2-}
Manganese (II)	Mn^{2+}	Sulfide	S^{2-}
Zinc	Zn^{2+}	Silicate	SiO_3^{2-}
Aluminum	Al^{3+}	Nitride	N^{3-}
Iron (III)	Fe^{3+}	Phosphate (V)	PO_4^{3-}

Steps to write a balanced equation

Step 1: Convert names of chemicals into their chemical formulae

Step 2: Balance the equation using the numbers

Step 3: Put in the state symbols

Soluble and Insoluble Substances in Water

Soluble substances – State symbol is usually (aq)	*Insoluble substances – State symbol is (s)*
All nitrate salts	-
Most chloride salts, except …	lead (II) chloride, silver chloride
Most sulfate salts, except …	lead (II) sulfate, barium sulfate, calcium sulfate
Carbonates of sodium, potassium and ammonium	…most other carbonates
Hydroxides and oxides of sodium, potassium and ammonium	…most other hydroxides and oxides

Common acids and alkali (used in aqueous form)

Acids – State symbol is(aq)	*Alkali – State symbol is(aq)*
Hydrochloric acid, HCl Nitric acid, HNO_3 Sulfuric acid, H_2SO_4	Sodium hydroxide, $NaOH$ Potassium hydroxide, KOH Aqueous ammonia, NH_3

*Note: All alkalis and acids dissolve in water. Bases do not dissolve in water.

Topic 6 – Stoichiometry and Mole Concept

What is a mole?

Definition: SI unit of substance. It is equal to the number of atoms in 12g of carbon-12. The number of atoms in a mole is approximately 6×10^{23}. This value is known as Avogadro's constant.

A chemist uses the quantity of a mole just like how 12 eggs are called a dozen eggs. Hence, he or she would use it when trying to find the amount of substance in a given mass of a certain substance.

Molar mass

Definition: The mass of one mole of any substance. It is numerically equal to the relative atomic mass/relative molecular mass of the substance expressed in grams.

Example

The A_r of carbon (C) = 12
$\therefore$, $molar\ mass\ of\ C = 12gmol^{-1}$

This means that every mole of carbon weighs 12g.

*Note: The M_r does not have a unit, as it is a relative quantity. However, the molar mass, despite being numerically the same as the M_r, has unit of g mol^{-1}, which means the mass of one mole of substance.

Finding the empirical formula of a substance from the % mass of each element within the substance

Step 1: Let the mass of the substance be 100g. Thus, you can determine the mass of each element within the substance according to its % mass. E.g.: If hydrogen is 2.4% in terms of the mass of compound X, in 100g of compound X, 2.4g will be hydrogen.

Step 2: Find the amount of substance in mol. E.g.: The molar mass of hydrogen is 1 g mol^{-1}. Hence, the no. of moles in 2.4g of hydrogen is $\frac{2.4g}{1g\ mol^{-1}} = 2.40\ mol$.

Step 3: Divide the substance by the smallest no. of mol. E.g.: In 100g of compound X, which is comprised of 39.0% sulfur, 58.6% oxygen and 2.4% hydrogen, the no. of moles of sulfur is 1.2 mol; the no. of moles of oxygen is 3.7 mol. As such, the smallest no. of mol is 1.2 mol. Hence, divide all the mol by 1.2 mol.

Step 4: Attain the simplest ratio of the different elements. E.g.: After dividing throughout, we find that the mol ratio of H:S:O = 2:1:3. Hence, the empirical formula of compound X is H_2SO_3.

*Note: If after dividing, even if you do not attain a whole number, your ratio should be a whole number. That means if you get 1.2:1, your ratio should be 5:4. However, if the number attained is very close to a whole number, you should round it off to the nearest whole number (e.g.: 1.05).

The molecular formula of the substance is obtained by dividing the M_r of the molecular formula by the M_r of the empirical formula, and multiplying the empirical formula through the ratio obtained through the division.

Percentage Composition by Mass

Definition of percentage composition by mass of a compound: The percentage that each element within the compound contributes to its mass.

Calculation steps

Step 1: Find M_r of compound

Step 2: Divide M_r of element within compound by M_r of compound itself.

*Note: If there is more than one atom of the same element within the compound, the A_r of the element is multiplied by the number of atoms within the element to obtain the M_r that will be used to determine the % composition by mass of a particular element.

Example:

In H_2O, if you want to find the percentage composition by mass of H in H_2O, the following steps are required.

Step 1: M_r of H_2O = $16 + 2\times1 = 18$

Step 2: % composition by mass = $\frac{2\times1}{18}\times100\% = 11.1\%$

Molecular mass and rate of diffusion

Gases can undergo effusion – they can pass through small, molecular-sized openings. Examples of effusion are air leaking out of balloons, basketballs and tyres.

*Note: Effusion is a type of diffusion.

In general, it is observed that denser gases diffuse slower than less dense gases. As density of a gas is directly proportional to the M_r of the gas, the general rule is that **the gas with a relatively larger M_r will have a slower rate of diffusion.**

OR

You can use this formula to understand the relationship.

Rate of diffusion of a gas = $\frac{k}{\sqrt{M_r}}$, where k is a constant.

Hence, the larger the gas particle, the slower the diffusion speed.

Limiting Reactant

Example

$N_2 + 3H_2 \rightarrow 2NH_3$

If you have 1 mole of N_2 and 4 mol of H_2, which is the limiting reactant?

When they are allowed to mix and react, in a theoretical complete reaction, 1 mol of N_2 and 3 mol of H_2 will react and be used up. Hence, 1 mol of H_2 is left over.

Explanation

The reactant, N_2, that is completely used up is called the limiting reactant. Hydrogen, in this example, is called the excess reactant, as an excess amount of H_2 is present initially. The amount of product, in this case NH_3 (ammonia), formed by the reaction is limited by the limiting reactant, N_2. As such, the addition of more excess reactant, in this case H_2, will have no effect on the amount of product, in this case ammonia, formed.

Molar Volume of Gases

Avogadro’s Law states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules.

Hence, at room temperature and pressure (r.t.p, 25°C and standard atmospheric pressure), it has been found that one mole of any gas occupies 24 dm^3. Hence, this amount is termed the **molar volume**.

Concentration of a Solution

Definition of a solution (YOU SHOULD KNOW THIS BY NOW): A homogeneous mixture containing a substance (solute) dissolved in another substance (solvent).

*Note: Water is an excellent medium for chemical reactions due to its solvent property (it can dissolve almost anything). A solution containing a substance dissolved in water is an aqueous mixture.

Saturated Solution

Definition: A solution that contains the maximum amount of solute that can be dissolved at a particular volume and temperature.

Two types of concentration

Mole concentration or Molarity

Definition: The number of moles of the solute dissolved in per unit volume (in dm^3) of the solution.

Units used: mol dm^{-3} or M (they mean the same thing)

The molarity of a solution can be calculated as such:

$$Molarity = \frac{No. of\ moles\ of\ solute\ (in\ mol)}{Volume\ of\ solution\ (in\ dm^3)}$$

Mass Concentration

Definition: The number of grams of the solute dissolved in per unit volume (in dm^3) of the solution.

Units used: g dm^{-3}

The mass concentration can be calculated as such:

$$Mass\ Concentration = \frac{No.\ of\ grams\ of\ solute\ (in\ g)}{Volume\ of\ solution\ (in\ dm^3)}$$

Relationship between Molarity and Mass Concentration

A formula to understand the relationship between molarity of a solution, its mass concentration and the molar mass of the solute is seen below:

$$Mass\ Concentration\ (g\ dm^{-3}) = Molarity\ (mol\ dm^{-3}) \times Molar\ mass\ of\ solute\ (g\ mol^{-1})$$

*Note: UNDERSTAND THE EQUATION BEFORE APPLYING!

Titration

Look at Notes – Volumetric Analysis: Titration Steps to find out how to titrate

Definition: A method of volumetric analysis in which the unknown concentration of a solution can be determined by reacting it with another solution of known concentration. This is done by measuring the exact volumes of the two reagents needed for a complete reaction.

Standard solution – Solution of known concentration

Acids and Bases

Chemistry of Acids

There are 4 chemical equations to remember for acids. They are:

1. Acid + Base → Salt + Water
   E.g.: HCl (aq) + $NaOH$ (aq) → $NaCl$ (aq) + H_2O (l)
2. Acid + Carbonate → Salt + Water + Carbon Dioxide
   E.g.: H_2SO_4 (aq) + $MgCO_3$ (s) → $MgSO_4$ (aq) + H_2O (l) + CO_2 (g)
3. Acid + Metals (some) → Salt + Hydrogen
   E.g.: $2HNO_3$ (aq) + Mg (s) → $Mg(NO_3)_2$ (aq) + H_2 (g)
4. Acid + Sulfites → Salt + Water + Sulfur Dioxide
   E.g.: HCl (aq) + Na_2SO_3 (aq) → $NaCl$ (aq) + H_2O (l) + SO_2 (g)

Chemistry of Bases

There are 4 chemical equations to remember for bases. They are:

1. Base + Acid → Salt + Water
   E.g.: $NaOH$ (aq) + HCl (aq) → $NaCl$ (aq) + H_2O (l)
2. Base + Ammonium Salt → Salt + Water + Ammonia
   E.g.: $NaOH$ (aq) + NH_4NO_3 (aq) → $NaNO_3$ (aq) + H_2O (l) + NH_3 (g)
3. Alkali + Salt solutions (some) → Salt + Insoluble hydroxide
   E.g.: $2NaOH$ (aq) + $CuSO_4$ (aq) → Na_2SO_4 (aq) + $Cu(OH)_2$ (s)

Classifying Acid

An acid can be classified in 3 different ways; concentration, strength and basicity.

Strength of an Acid

Definition: Degree of dissociation or ionisation.

A strong acid is one that undergoes almost 100% dissociation, and is fully ionised.

A weak acid experiences partial dissociation (e.g.: 75%, 50%), and is only partially ionised.

*Note: At any point in time, molecules are constantly dissociating and joining again, but the net dissociation is the % dissociation above.

Concentration of an Acid

Definition: Amount of substance dissolved in solvent

Hence, by comparison, an acid of concentration 0.100 mol dm^{-3} is considered dilute while an acid of concentration 2.00 mol dm^{-3} is considered concentrated.

Basicity

Acids are classified as monobasic/monoprotic, dibasic/diprotic, and tribasic/triprotic, according to the number of H atoms they possess that can ionise to form H^+ ions (Note: They actually form H_3O^+).

For example:

HCl is a monobasic acid, as it has only 1 H atom that can ionise to form H^+ ions.

CH_3COOH is monobasic as well, as it has only 1 H atom that can ionise to form H^+ ions. (The other 3 H atoms cannot dissociate. Note that in organic acids, USUALLY only the last H in the chemical formula can ionise.

H_3PO_4 is tribasic, as it has 3 H atoms that can ionise to form H^+ ions.

*Note: Increased basicity does not necessarily affect pH, as even though tribasic acids produce more H^+ ions per molecule when compared to monobasic or dibasic acids, these monobasic and dibasic acids might be stronger, which means that the ionisation in such acids is greater than that in tribasic acids.

Example: $(COOH)_2$ vs HCl

Oxalic acid is a dibasic acid. However, as oxalic acid is a weak acid, it will dissociate far less than HCl. As such, at the same concentration, HCl will have a lower pH than $(COOH)_2$.

Oxides

There are multiple types of oxides, and in this subtopic, we will be covering the classification of these oxides.

*Note: Refer to Notes – Acids and Bases: Types of Oxides for the flowchart.

Metal Oxides

There are 2 types of metal oxides: basic and amphoteric.

Basic Oxides

Definition: Reacts with an acid to form a salt

Examples: Na_2O, CaO, CuO

Amphoteric Oxides

Definition: Reacts with both an acid and a base to form a salt.

ONLY 3 AMPHOTERIC OXIDES ARE PRESENT; ZnO, PbO, Al_2O_3

*Note: Remember Z.A.P.

Amphoteric oxides, when reacting with a base, form a complex salt.

E.g.: $Al_2O_3\ (s) + 2NaOH\ (aq) \rightarrow 2NaAlO_2\ (aq) + H_2O\ (l)$

Important to note: Hydroxides behave like oxides. Therefore, $Al(OH)_3$, $Pb(OH)_2$ and $Zn(OH)_2$ are amphoteric as well.

pH of a Solution

Definition: Measure of concentration of H^+ ions (technically H_3O^+).

pH = $-\lg[H^+]$

*Note: A change of 1 unit of pH changes the H^+ ion concentration by a factor of 10.

pH can extend beyond the 0-14 range in VERY concentrated acids and bases.

An aqueous solution always contains H^+ and OH^- ions, due to the small extent of partial dissociation of water molecules. The equation is as follows:

$H_2O \rightleftharpoons H^+ + OH^-$

Solution	Relative concentration of Ions	pH
Acidic	$[H^+] > [OH^-]$	pH < 7
Alkaline	$[H^+] < [OH^-]$	pH > 7
Neutral	$[H^+] = [OH^-]$	pH = 7

*Note: Many students think that an acidic solution only contains H^+ and no OH^- ions. This is untrue. In an acidic solution the concentration of H^+ is greater than the concentration of OH^-, that is why the pH is less than 7.

Measuring pH

pH can be measured using chemical indicators that show different colours at different pH values, e.g.: Universal Indicator and litmus paper.

A few drops of indicator solution are added to the solution to be tested, and the resultant solution colour is then compared with the colour chart and the corresponding pH value.

It can also be measured using a pH meter/sensor.

pOH

Definition: Measure of concentration of OH^- ions.

In general, $pOH = 14 - pH$.

*Models of Acids and Bases

Arrhenius' Model

Definitions

Definition of an acid: Any substance that ionises when it dissolves in water to give the H^+ ion.

Definition of a base: Any substance that gives the OH^-, or hydroxide, ion when it dissolves in water.

What does this model explain?

This model explains why all acids have similar properties to one another and why bases have similar properties to one another as well.

It also explains the observation that acids and bases counteract one another. This idea, that a base can make an acid weaker, and vice versa, is called neutralisation.

What can it not explain?

The model can only be applied to reactions that occur in water because it defines acid and bases in terms of what happens when compounds dissolve in water. It does not explain reactions such as those that occur in the gaseous phase (e.g.: $HCl (g) + NH_3 (g) \rightarrow NH_4Cl$).

This model also cannot explain why compounds such as Na_2CO_3 and NH_3 have characteristic properties of bases as only compounds that contain the OH^- ion can be classified as Arrhenius bases.

The existence of the H^+ ion in solution was also cast in doubt as it is so tiny and the electric field it generates in its surroundings is likely so intense that it attracts any molecule with any unshared pair of electrons, such as H_2O.

A hydronium/oxonium ion if formed when the H^+ ion reacts with a water molecule.

Chemical Equation: $H^+ + H_2O \rightarrow H_3O^+$

Brønsted-Lowry Model

Definitions

Definition of an acid: A substance which can donate a hydrogen ion/proton to another substance. (Note that H^+ is actually just a proton because H only has one electron)

Example: The dissociation of HCl in water is not actually dissociation, but rather a donation of a H^+ ion from HCl to a H_2O molecule to form H_3O^+.

Chemical Equation: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

*Note: pH is technically a measure of the concentration of H_3O^+ ions as H^+ ions are too reactive to exist and hence in water bonds with the unshared electron pairs on the oxygen molecule of H_2O to form H_3O^+ ions.

This model not only encompasses all acids that fit into the Arrhenius definition but also include acid-base reactions that occur in the gas phase, like the one mentioned before between HCl and NH_3.

Definition of a base: A substance which can accept a hydrogen ion/proton from another substance.

This definition is very different from the Arrhenius definition of a base, and explains why substances like NH_3 and Na_2CO_3, which do not contain the OH^- ion, display characteristic properties of a base.

Example: Baking soda acts as a base by accepting a H^+ ion from an acid as shown in the equation below.

$HCO_3^- + H^+ \rightarrow H_2CO_3$ (unstable) $\rightarrow H_2O + CO_2$

Carbonates also react in the same way:

$CO_3^{2-} + 2H^+ \rightarrow H_2CO_3$ (unstable) $\rightarrow H_2O + CO_2$

Ammonia also fits the definition by reacting in a similar way:

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

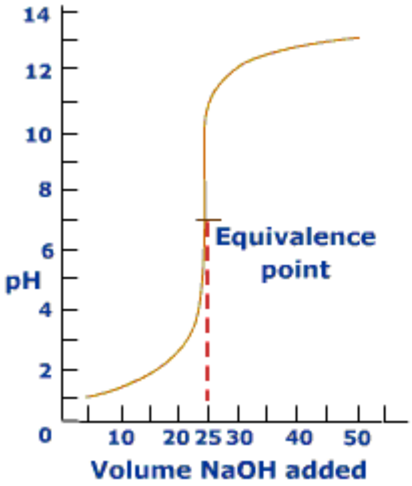
Insoluble metal oxides are also bases according to the definition, because the O^{2-} ion in these oxides accept H^+ ions from acids to form water. An example is MgO:

$MgO\ (s) + 2HCl\ (aq) \rightarrow H_2O\ (l) + MgCl_2\ (aq)$

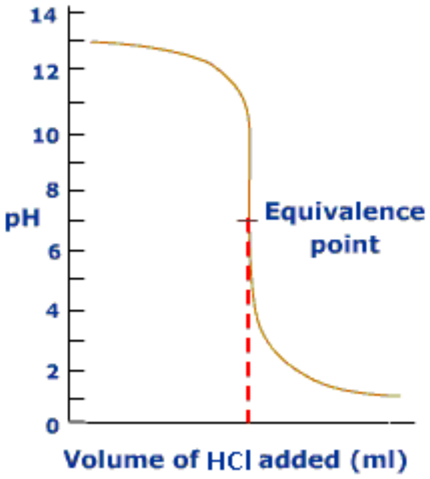
Lewis Model

Acid-Base Titration Curves

Strong Acid-Strong Base Curves



Titration curve of strong acid (HCl) with a strong base (NaOH)



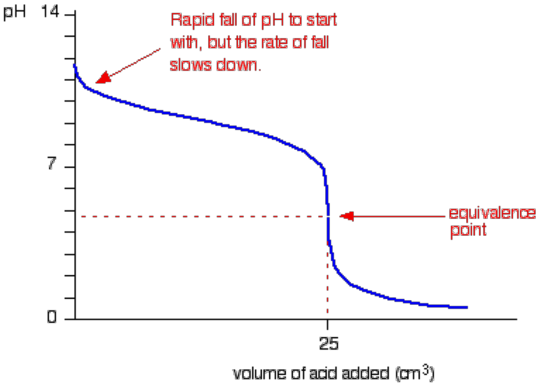
Titration curve of strong base (NaOH) with strong acid (HCl)

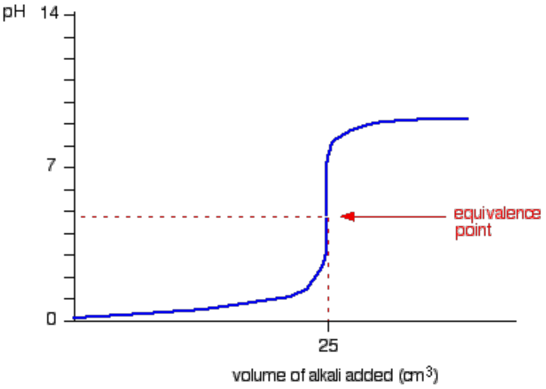
Initially, pH decreases slowly because $[H^+]/[OH^-]$ (concentration is represented by the square brackets) is still very high.

Then, there is a very steep drop/gain because one drop added contains a very high concentration of H^+/OH^-.

Equivalence point = pH 7

Strong Acid-Weak Base Titration



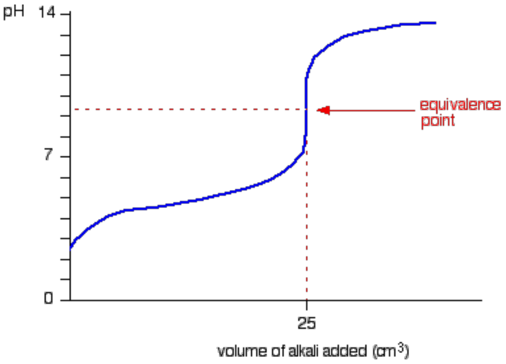


*Note: The gentle slope initially is called the buffer region. This is because a buffer solution is set up, and so pH falls slowly. A buffer solution sets up an equilibrium between the acid and the conjugate base (Definition: Removal of proton/H+ ion from an acid), and so the equilibrium is shifted slightly to the left, thus resulting in a smaller-than-expected change. (You DO NOT need to know this)

Equivalence point = pH 5 (approximately)

Acidic salt is formed.

Weak Acid-Strong Base Titration



*Note: when a weak acid is added to a strong alkali, the graph is laterally inverted.

The initial decrease in the rate of increase in pH is due to the formation of a buffer solution.

Equivalence point = pH 9 (approximately)

There is alkali salt formed.

General Patterns

In general, when titrated:

- Strong acid + Strong base – neutral solution
- Strong acid + Weak base – acidic solution (pH < 7)
- Weak acid + Strong base – Alkali solution (pH > 7)

Observations of Reactions Involving Acids and Bases

4 types of observations must be recorded. They are:

1. Production of a gas
2. Production of a solid (precipitate)
3. Dissolving of solid to form a solution
4. Colour change of pH indicator

*Note: In all observations, state and colour MUST be mentioned.

Rules of Solubility

The following are the rules of solubility:

1. Chlorides, Bromides, and Iodides – Most are soluble in water. The two common exceptions are: lead (II) and silver chloride/bromide/iodide.
2. Sulfates – Most are soluble in water. The three common exceptions are barium sulfate, lead (II) sulfate, and calcium sulfate (sparingly soluble).
3. Carbonates – Most are insoluble in water. The two exceptions are ammonium carbonate and Group I carbonates.
4. Metal Oxides and Hydroxides – Most are insoluble. Exceptions are ammonium oxide/hydroxide, Group I oxides/hydroxides. Calcium oxide/hydroxide is sparingly soluble.
5. Group I and Ammonium Compounds – All are soluble.
6. Nitrates – All are soluble.

Steps to apply the rule:

Step 1: Check if substance is a nitrate. If it is, then it is soluble.

Step 2: Check if substance is a Group I or ammonium compound. If it is, then it is soluble.

Step 3: Identify type of substance. (e.g.: if it is a chloride, sulfate, etc.)

Step 4: Apply rules 1-4 if Step 1 and 2 have not identified if it is soluble.

Precipitation Reactions

Definitions

Precipitation: Formation of a solid when 2 aqueous susbtances are mixed. An insoluble substance would appear as a precipitate when it is formed.

Formation of a precipitate requires that one of the reactants must contain the cation and the other must contain the anion that forms the precipitate.

Example: To form AgCl as a precipitate, one reactant must contain Ag^+ and the other Cl^-.

Chemical Equation: $AgNO_3\ (aq) + NaCl\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$

Collection of precipitate

The steps to collecting the precipitate are:

1. Filter the mixture
2. Wash the residue on the filter paper with distilled water
3. Dry the solid by pressing it between sheets of filter paper or in a low-heat oven.

Ionic Equations

When substances react, not all the particles (atoms/ions/molecules) take part in the reaction.

Definitions

Ionic Equation: An equation the specifically shows which atoms/ions/molecules take part in a reaction.

Spectator Ions: Free ions present at the start and at the end of a reaction. They remain unchanged by the reaction.

When writing an ionic equation from a chemical equation, the aim is to eliminate spectator ions.

Steps

1. Write the balanced chemical equation.
   E.g.: $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
2. Put in state symbols. (Apply solubility rules!)
   E.g.: $AgNO_3\ (aq) + NaCl\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$
3. Write out the free ions in aqueous solution.
   E.g.: $Ag^+\ (aq) + NO_3^-\ (aq) + Na^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s) + Na^+ + NO_3^-\ (aq)$
4. Cancel away corresponding ions on both the left and right side of the equation.
   E.g.: $Ag^+\ (aq) + \cancel{NO_3}^-\ (aq) + \cancel{Na}^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s) + \cancel{Na}^+\ (aq) + \cancel{NO_3}^-\ (aq)$
5. Writing down the formulas that have not been cancelled gives the ionic equation.
   E.g.: $Ag^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s)$

*Note: The net charge on both sides must be balanced.

Neutralisation (Acid-Alkali) Reactions

The ionic equation for all neutralisation equations is the same as the soluble salt formed remains as ions in the solution.

Ionic Equation: H^+ (aq) + OH^- (aq) → H_2O (l)

Precipitation Reactions

The ionic equation for all precipitation reactions involve ONLY the ions of the precipitate.

Example: $AgNO_3$ and NaCl forming a precipitate.

Chemical Equation: $AgNO_3$ (aq) + NaCl (aq) → AgCl (s) + $NaNO_3$ (aq)

Ionic Equation: Ag^+ (aq) + Cl^- (aq) → AgCl (s)

Acid + Metal

The ionic equation only involves the metal and the H^+ ion from the acid.

E.g.: Zn and H_2SO_4

Chemical Equation: Zn (aq) + H_2SO_4 (aq) → $ZnSO_4$ (aq) + H_2 (g)

Ionic Equation: Zn (aq) + $2H^+$ (aq) → Zn^{2+} (aq) + H_2 (g)

Acid + Soluble Carbonate

The ionic equation involves only H^+ ions and CO_3^{2-} ions.

E.g.: HNO_3 and Na_2CO_3

Chemical Equation: $2HNO_3$ (aq) + Na_2CO_3 (aq) → $2NaNO_3$ (aq) + H_2O (l) + CO_2 (g)

Ionic Equation: $2H^+$ (aq) + CO_3^{2-} (aq) → H_2O (l) + CO_2 (g)

Displacement Reaction

The ionic equation involves only the ion that is reduced and oxidised.

Chemical Equation: Zn (s) + H_2SO_4 (aq) → Cu (s) + $ZnSO_4$ (aq)

Ionic Equation: Zn (s) + Cu^{2+} (aq) → Zn^{2+} (aq) + Cu (s)

Redox Reactions

Definitions

Oxidation: Process when a substance's oxidation state increases through the loss of electrons.

Reduction: Process when a susbtance's oxidation state increases through the gain of electrons.

Oxidation State: Arbitrary hypothetical number given to atom would have if all bonds to atoms of different elements were 100% ionic. (Definition off of Wikipedia☺)

Half Equations

Reactions involving a transfer of electrons can be viewed as 2 halves. One half equation shows oxidation while the other shows reduction.

Consider the redox reaction:

$2Na + Cl_2 \rightarrow 2NaCl$

($NaCl$ comprises Na^+ and Cl^-)

The 2 half equations are:

$Na \rightarrow Na^+ + e^-$ (Na lost an electron, hence Na is oxidised)

$Cl_2 \rightarrow 2Cl^- + 2e^-$ (Cl_2 gained 2 electrons, hence Cl_2 is reduced)

2 half equations can be combined to form the full chemical or ionic equation.

*Note: The number of electrons lost or gained are the same; if it is not, multiply the half equation by a suitable number.

Oxidation State/Number

Rules for assigning oxidation states:

Rule 1: Elements uncombined with other elements are of oxidation state 0. (E.g.: Oxidation state of N in N_2 is 0)

Rule 2: In simple ions, the oxidation state is the charge on the ion. (E.g.: Oxidation state of O^{2-} is -2)

Rule 3: In polyatomic ions, the sum of the oxidation states of the individual elements that constitute the polyatomic ions is equal to the charge of the ion itself.

Rule 4: The oxidation state of hydrogen in all its compounds is +1 except in metal hydrides where its oxidation state is -1.

Rule 5: The oxidation state of oxygen in all its compounds is -2, except in peroxides where its oxidation state is -1.

Rule 6: The oxidation state in Group I elements in their compounds is +1, for Group II elements in their compounds is +2, and for aluminium in its compounds is +3.

Rule 7: There are many oxidation states for Group VII elements in their compounds but the usual one is -1.

Rule 8: The sum of the oxidation states of all the elements in a molecule or compound is 0.

Reducing and Oxidising Agents

Definitions

Oxidising Agents: Substances that cause others to undergo oxidation. They themselves are reduced.

Reducing Agents: Substances that cause others to undergo reduction. They themselves are oxidised.

Laboratory Tests for Reducing or Oxidising Agents

Reactants that undergo reduction or oxidation with observable colour changes are used to test for the oxidising and reducing properties of other substances.

The chemicals to test for the presence of reducing agents are $KMnO_4$ and $K_2Cr_2O_7$. A reagent used to test for oxidising agents is KI.

Test for reducing agents using $KMnO_4$

	MnO_4^-	Mn^{2+}
Oxidation state of Mn	+7	+2
Colour	Purple	Colourless

Preparation of reagent: 1 cm^3 of $KMnO_4$ solution into a test tube, followed by 1 cm^3 of dilute sulphuric acid.

Test procedure: A few drops of the acidified $KMnO_4$ is added to the unknown solution.

Observation and Conclusion:

In the purple MnO_4^- turns colourless, then MnO_4^- has been reduced to Mn^{2+}. This means the unknown substance caused the reduction and hence contains a reducing agent.

Test for reducing agents using $K_2Cr_2O_7$

Dichromate (VI) ions can undergo reduction to chromium ions.

	$Cr_2O_7^{2-}$	Cr^{3+}
Oxidation state of Cr	+6	+3
Colour	Orange	Green

Preparation of reagent: 1 cm^3 of $K_2Cr_2O_7$ solution is poured into a test tube, followed by 1 cm^{3+} of dilute sulfuric acid.

Test procedure: A few drops of the acidified $K_2Cr_2O_7$ is added to the unknown substance (solution).

Observation and Conclusion: If the orange $Cr_2O_7^{2-}$ solution turns green, then it has been reduced to Cr^{3+}. This means the unknown substance caused the reduction and hence contains a reducing agent.

Test for oxidising agents using KI

Iodide ions can be oxidised to iodine.

	I^-	I_2
Oxidation state of I	-1	0
Colour	Colourless	Brown

Test procedure: A few drops of aqueous KI is added to the unknown substance.

Observation and Conclusion: If the colourless I^- solution turns brown, then I^- has been oxidised to I_2. This means the unknown substance caused the oxidation and hence contained an oxidising agent.

Chemical Energetics

There are 3 important types of energy conversions:

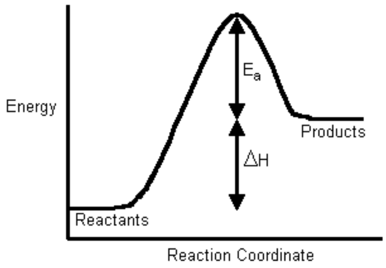
1. Chemical Reactions that involve a heat/light change (exothermic and endothermic reactions)
2. Chemical Reactions that involve an electrical change (in cells and batteries; in electrolysis)
3. Nuclear reactions (nuclear fission)

In this topic, we will be focussing on the 1st type of energy conversion.

Definitions

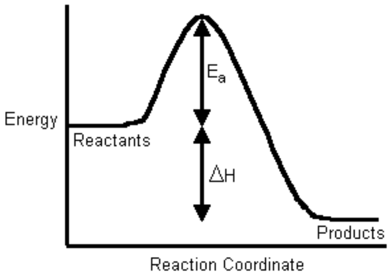
ΔH/Enthalpy change: Net change in heat of reaction

Exothermic Reaction: A reaction in which energy (heat) is given out. This causes a temperature rise in surroundings.



The energy content of the products is lower than the energy content of the reactants. Hence, the ΔH is negative.

Endothermic Reaction: A reaction in which energy (heat) is taken in. This causes a temperature drop in surroundings.



The energy content of the products is higher than the energy content of the reactants. Hence, the ΔH is positive.

Activation Energy (E_A): The initial energy necessary to kick-start the reaction, so that bonds can be broken and formed.

Bond-Breaking

Bond-breaking is endothermic as it requires heat to overcome the forces of attraction between particles.

Bond-Making

Bond-making is exothermic as the potential energy stored within bonds is negative, and so when bonds are formed, there is a reduction in ΔH.

Calculation of ΔH

Calculating ΔH requires the bond energies of the bonds between the particles in the reactants.

Definitions

Bond Energy: Energy absorbed in breaking a covalent bond.

Calculation

Bond breaking is endothermic. Hence, all the bond energies involved in bond breaking is given a '+' sign (that is, it is taken as positive).

Bond making is exothermic. Hence, all the bond energies involved in bond making is given a '-' sign (that is, it is taken as negative).

Hence, the equation to calculate ΔH is:

$$\Delta H = Total\ energy\ of\ bond\ breaking - Total\ energy\ of\ bond\ making$$

E.g.: Hydrogen (H_2) and Chlorine (Cl_2)

Chemical Equation: $H_2 + Cl_2 \rightarrow 2HCl$

Bond energy of H-H: 436 kJ mol^{-1}
Bond energy of Cl-Cl: 242 kJ mol^{-1}
Bond energy of H-Cl: 431 kJ mol^{-1}

Hence,

$$\Delta H = (energy\ to\ break\ H - H\ bond) + (energy\ to\ break\ Cl - Cl\ bond) \\ - (energy\ released\ whn\ two\ H - Cl\ are\ formed)$$

$$\Delta H = 436 + 242 - (431 \times 2)$$

$$\Delta H = -184kJ$$

As ΔH is negative, the compound formed is an exothermic compound. It has a low energy content, which means it is energetically stable (with respect to its elements). Hence, the bonds between the atoms are strong.

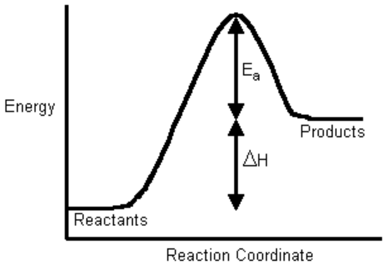
Reaction Kinetics

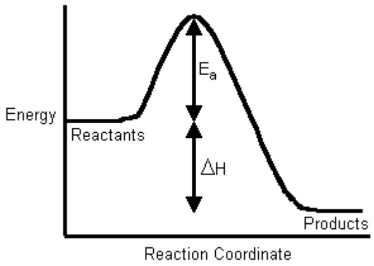
For a chemical reaction to occur between reactants, effective collisions must occur between the reactant particles.

Definitions

Effective Collisions: Collisions between reactant particles in such a way and with such force that bond breaking and bond making occur.

Activation Energy: Minimum amount of energy before a reaction will occur.





The energy diagrams above show that activation energy (E_A) is an “energy barrier”/”energy hill” over which reactants must overcome before they can react.

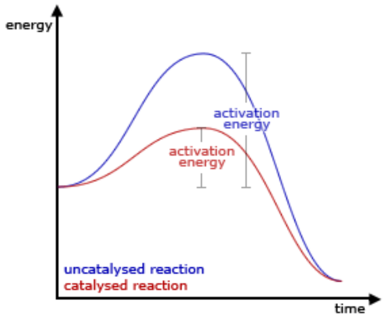
*Note: When E_a is high, it suggests that the reaction is slow, as only a few of the collisions give particles sufficient energy to produce a reaction. The converse is true, that is, when the E_a is low, the reaction is faster as many of the collision give particles sufficient energy to produce a reaction.

5 Factors that Increase Rate of Reaction

1. Increase in Temperature:
   At a higher temperature, the particles have more energy. Hence, they move faster, colliding harder with one another, thus increasing rate of reaction. This is possible because of an increase in the number of reactant particles having the activation energy.
2. Increase in concentration of reactants:
   At a higher concentration, there are more particles per unit volume. Hence, particles are closer to one another, and the frequency of collisions increases, hence increasing reaction rate.
3. Increase pressure (for gaseous reactant)
   At higher pressure, the particles are in closer proximity to one another, thus colliding more frequently. This results in increased reaction rate.
4. Increase surface area of reactants
   An increased surface area means that the area of contact between the reactants increases, hence the number of collisions per unit time also increases. E.g.: Magnesium ribbon, magnesium powder
5. Presence of catalysts
   A catalyst provides an alternative pathway which has a lower activation energy. The particles of the catalyst somehow affect the collision of the reactant particles, usually by providing a

site for collision to occur. Hence, it is said that a catalyst provides an alternative route of reaction.

*Note: Increased amount of catalyst will also increase the rate of reaction, though a small amount will already have a large effect.



The above diagram compares the activation energy of an uncatalysed reaction to a catalysed one. Note that the activation energy is not reduced, but rather there is an alternative activation energy that has lower activation energy.

*Note: Catalysts do take part in the reaction, but it itself is not chemically reacted at the end of the reaction.

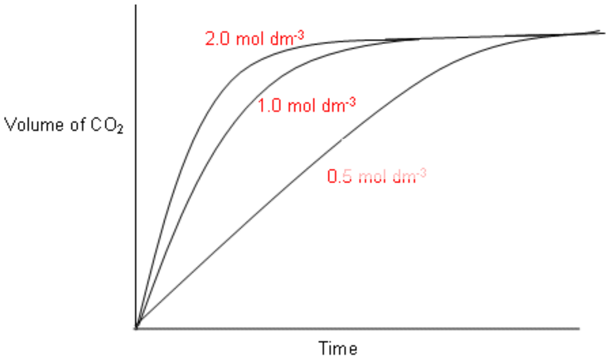
4-Part Answering Guide

1. Factor and what it does
2. Hence, rate of reaction increases due to…
3. Increase in frequency of collisions
4. Hence, number of effective collisions increases
5. Between …. (reactant 1) and …. (reactant 2)

Measuring rate of reaction

One way is to measure the decrease in the mass of the reaction mixture at regular intervals. The gradient of the curve of the graph plotted from the decrease in mass will tell you the rate of reaction.

Another is to measure the increase in the volume of gas formed. The gradient will again tell you the rate of reaction.



Note that the reaction rate is fastest at the start of the reaction due to the fact that the reactants are at the highest concentration.

The gradient becomes gentler over time, hence suggesting that the rate of reaction decreases with time. (It doesn't matter if the gradient is negative or not)

*Note: Monitoring the rate of reaction in this manner means that one of the products must be a gas. Other methods are:

1. Measuring changes in amount of reactant or product by titration. Small amounts of the reaction mixture is removed at fixed intervals and titrated to check the amount of a reactant or product.
2. Measuring the changes in intensity of colour of reaction mixture. One of the reactants or products must be a coloured substance, and the intensity of colour can be followed using a colourimeter.
3. Measuring the changes in pressure in a closed reaction vessel. This is suitable for reactions in the gas phase, and where the number of gaseous products differ from the number of moles of gaseous reactants.
4. Measuring the temperature change as the reaction proceeds.

*Catalysts