End of Year Notes: Chemistry

Topic 1: Nature and Classification of Matter

Definitions

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.

Similarities & Differences between Elements and Compounds

Similarities	Differences
Both are made up of atoms as the building blocks	An element is made up of only one type of atom while a compound is made up of more than one type of atom
Both are pure substances	An element can be metallic or non metallic, while a compound is formed by combining a metallic element with a non metallic element, or several non metallic elements together
Both elements and compounds have fixed melting and boiling points	An element cannot be chemically split into simpler substances, while a compound can be decomposed into elements or simpler compounds.

Similarities and Differences between Compounds and Mixtures

Similarities	Differences
Both are made up of more than one type of atom	A mixture can be separated into its components by physical methods, while a compound can be separated into its elements (or simpler compounds) only by chemical methods
	A mixture has no fixed melting or boiling point while a compound has fixed melting and boiling points
	No chemical changes take place when a mixture is formed from its constituents but a chemical change occurs when a compound is formed
	The proportion of constituents in a mixture can vary but the proportion of elements in a compound is fixed.
	The properties of a mixture are the same as those of its constituents but the properties of a compound differ from those of its constituent elements.

Hierarchy:

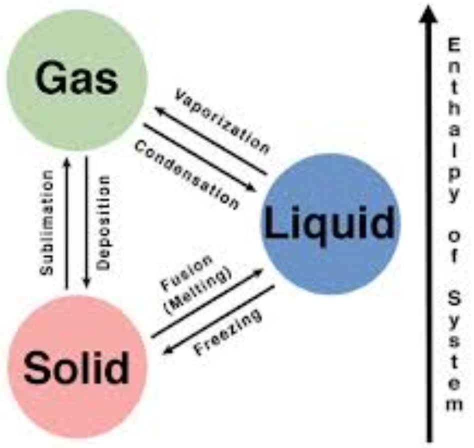
- Pure substances
    - Elements
        - Metals (draw the staircase from boron's left and bottom, note: hydrogen is an exception – it is not metallic)
        - Non metals
    - Compounds
- Mixtures
    - Element + Element
    - Element + Compound
    - Two or more compounds

Kinetic Theory

- All matter is made up of atoms (the smallest unique particle of each element). A gas particle could be an atom or a molecule.
- Atoms/molecules have energy of motion that is felt as temperature. The motion of atoms/molecules can be in the form of vibration or linear motion.
- The pressure of a gas is due to the motion of the atoms/molecules of gas striking the object bearing the pressure
- There is a very large distance between the molecules of a gas (such that the size of the gas atom/molecule is considered negligible).

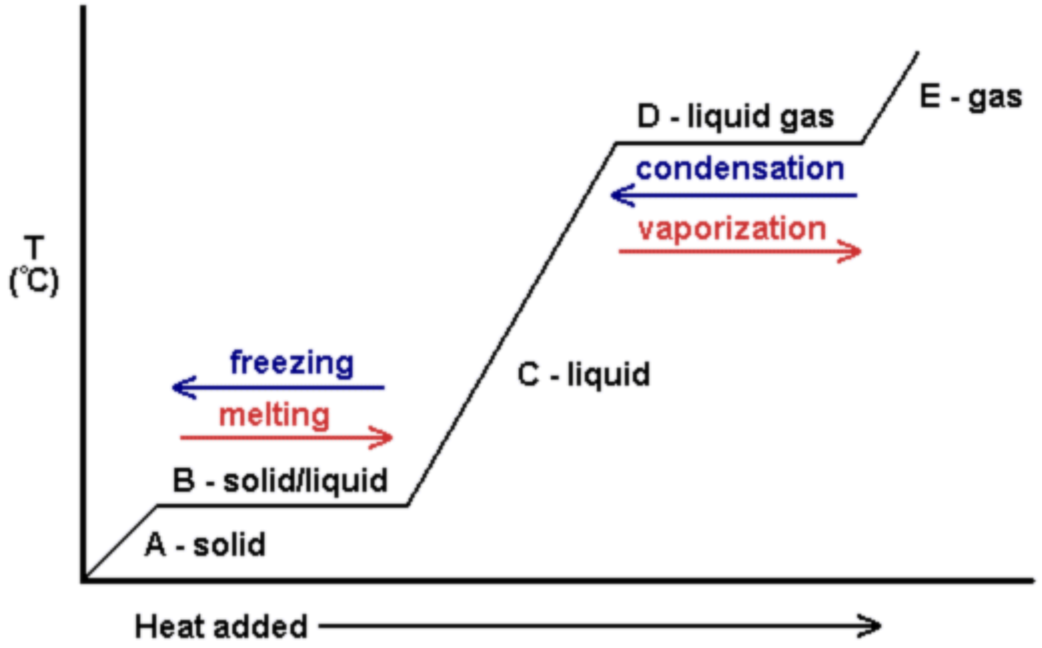
Particulate Nature of Matter

	Solid	Liquid	Gas
Energy Content	Low	Moderate	High
Arrangement of Particles	Close together in an orderly arrangement	Close together but no orderly arrangement	Far apart and no orderly arrangement
Motion	Vibration but no translation (cannot move on their own)	Vibration and Translation	Vibration and Translation
Shape and Volume	Fixed shape and fixed volume (cannot be compressed)	Fixed volume but no fixed shape (Cannot be compressed)	No fixed shape, and no fixed volume (can be compressed, and takes the volume of the container)
Forces of Attraction between particles	Strong	Moderate	Weak or none



Note: When explaining particulate nature, use ADAM (Arrangement, Distance between Particle, Attractive Forces, Movement of particles) to describe solids liquids or gases.

Heat curve:



The heat curve above shows the change in energy (in calories) and temperature of a liquid (for example, water) as it undergoes a phase transition.

At 100°C the liquid water will begin to undergo a phase transition into the gas phase. Energy introduced into the liquid will not go into increasing the temperature – it will be used to send molecules of water into the gas state by breaking the forces of attraction between the particles. The water will remain at 100oC until all of the water has finished transition to the gaseous state.

Evidence for particulate nature of matter:

- Diffusion: The movement of particles from an area of high concentration to an area of low concentration

  - Temperature: The higher the temperature, the faster the particles diffuse.
  - Mass of particles: The greater the mass of particles, the slower the particles diffuse.

Misconceptions:

- When you mix 50ml of water with 50ml of ethanol, you will not get a mixture of 100ml. This is because the particles of water filled up the spaces found between the particles of ethanol before contributing to increasing the volume of the mixture.
- Matter may be coloured (for example, sulfur powder is yellow) but the particles themselves are not coloured.
- Substances may be hot or cold, but the particles are not hot or cold. The particles vibrate fast in hot substances and slow in cool substances.
- Matter expands when heated, but the particles do not expand – it is the spaces between the particles that expand.

Topic 2: Experimental Design and Purification Techniques

Gases

Collection:

1. Displacement of water
   a. The gas is passed through a delivery tube which enters the bottom of a gas jar filled with water. As water enters the gas jar, water is pushed out (thus, displacement).
   b. Note: Gas must not be soluble in water.
   c. Example: Hydrogen, methane
2. Downward delivery
   a. Gas is passed through a delivery tube into a gas jar.
   b. Note: Gas must be denser than air.
   c. Example: Carbon dioxide, Bromide, Chloride
3. Upward delivery
   a. Gas is passed through a delivery tube into a gas jar.
   b. Note: Gas must be less dense than air.
   c. Example: Helium

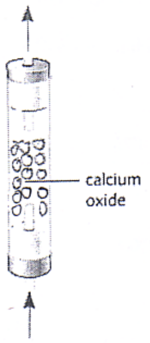
Measurement:

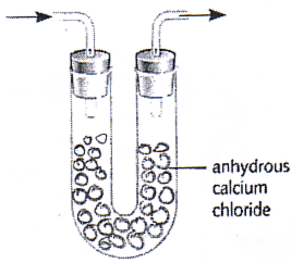
- A gas is measured by pumping the gas produced through a gas syringe. The plunger of the syringe will be pushed out to allow us to read off the volume of the gas.

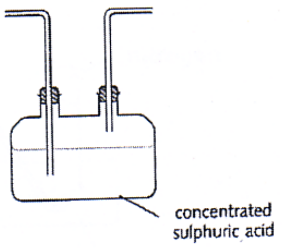
Drying Agents:

- A gas produced is often contaminated with water vapour. Thus the gas is passed through a drying agent to dry the gas. There are 3 types of drying agents:
  - Basic/Alkaline drying agents (E.g. CaO, calcium oxide) – Cannot dry acidic gases, such as CO_2
  - Acidic drying agents (E.g. H_2SO_4, sulfuric acid, or anhydrous copper II oxide) – Cannot dry basic/alkaline gases, such as NH_3 (ammonia)

        - Anhydrous copper II oxide is also used to test for the presence of water vapour (will turn blue from white if water vapour is present).
    - Neutral drying agents (E.g. $CaCl_2$, calcium chloride) – Can dry neutral, acidic, and basic gases
- Set ups for drying gases:







Liquids

Common apparatus:

- Measuring cylinder
    - 0 decimal places
    - Not a fixed volume
    - Note: You cannot state the volume as 5.0ml because it can only read to an accuracy of 0dp.
- Burette
    - 2 decimal places (the second one is a 5 or 0)
    - Not a fixed volume
- Pipette
    - 1 decimal place
    - Only for fixed volumes (25ml or 50ml)
- NEVER use a beaker to measure the volume of a liquid.

Purification Techniques

1. Filtering and Evaporating
    a. Filtering – **To remove insoluble impurities from a substance**
    b. Evaporating – **To obtain the solute from a solution of a soluble impurity and a solvent**
2. Crystallization
    a. Crystals contain the water of crystallization (and so excessive heating would drive off the water and prevent crystals from forming). General rule of thumb is to heat solution until 1/3 of the original solution remains (this is when it would be saturated).
    b. Note: Crystallization makes use of the solubility curve, which is why the crystals form due to oversaturation as the temperature goes down.
3. Sublimation

    a. Some substances sublime directly from the solid state to the gaseous state (eg. Iodine, carbon dioxide, ammonium chloride) – so **sublimation can be used to separate the components of a mixture in which one substance sublimes and the other does not.**

4. Separating Funnel
    a. **A separating funnel is used to separate 2 immiscible liquids.**
5. Simple Distillation
    a. **Simple distillation is used to recover the solvent from a solution** sd das
    b. Boiling chips – ensure smooth boiling by allowing the formation of smaller air bubbles
    c. Thermometer placed at exit of flask – to measure the temperature of the pure vapour exiting the flask to determine the boiling point of the solvent
    d. Cooling water from the bottom of the condenser – to ensure that the bottom of the condenser is the coolest part and that any vapour that did not condense initially will condense now, and also to ensure that the whole condenser is filled with cool water.
6. Fractional distillation
    a. **Fractional distillation is used to separate a mixture of miscible liquids with different boiling points (eg. Water and ethanol)**
    b. Bunsen burner is not used as some immiscible liquids might be flammable (such as ethanol) so a boiling water bath is used instead.
    c. Tiny glass beads are used in the fractionating column to increase the surface area to ensure that the vapour condenses repeatedly to ensure that impure vapours formed by evaporation do not escape.
    d. The thermometer would first record the temperature of the miscible liquid with the lower boiling point, stagnate there, and then increase to the boiling point of the next miscible liquid, then stagnate there, etc.
7. Reverse Osmosis
    a. **Reverse osmosis is used to purify impure water**
    b. Impure water has pressure applied to it and is passed through a partially permeable membrane which has tiny holes or pores.

Substance Purity

- An pure substance has fixed boiling and melting points
- The presence of an impurity increases the boiling point and decreases the melting point, and boiling and melting no longer occurs at a fixed temperatures (but rather, over a temperature range).
- The impurity can also be measured with the RF (paper chromatography), by measuring the distance travelled by solute / distance travelled by solvent.

Gas Tests

- Chlorine
    - Test: Test the evolved gas with moist blue litmus paper
    - Result: Blue litmus paper turns red and then bleached
- Hydrogen

  - Test: Insert a lighted wooden split into the test tube with gas sample.
  - Result: Gas extinguishes the lighted split with a pop
- Oxygen
  - Test: Insert a glowing split into the test tube with gas sample.
  - Result: The gas rekindled the glowing splint.
- Sulfur Dioxide
  - Test: Dip a piece of filter paper into acidified potassium dichromate (VI), and hold the filter paper at the mouth of the test tube with gas sample
  - Result: The orange potassium dichromate (VI) turns green.
- Carbon Dioxide
  - Test: Pass the gas through limewater
  - Result: A white precipitate is formed
- Ammonia
  - Test: test the gas with a piece of moist red litmus paper
  - Result: The red litmus paper turns blue

Topic 3: Atomic Structure

Bohr Model

A positively charged nucleus (containing protons and neutrons) are surrounded by electrons that travel in circular orbits or "shells" around the nucleus

- Electrons can only trave,l in special energy levels ("shells"), at certain set distances away from the nucleus with specific energies. Electrons do not continuously lose energy as they travel. Electrons only gain and lose energy by jumping from one allowed energy level/shell to another.
- The maximum number of electrons that can be accommodated in any shell is $2N^2$ where N equals the number of the energy level (shell level).
  - Therefore maximum number of electrons in shell 3, 4 and 5 are: 9x2, 16x2, 25x2
  - Few exceptions though (such as iodine and potassium, which is 2.8.8.1 instead of 2.8.9)
  - Atoms often tend to gain or lose electrons to attain a noble gas configuration, thus making themselves ions.

Electron Cloud Model (Wave Mechanical Model)

- It was discovered that electrons have both particle like properties and wave like properties (thus the positions of the electrons are unknown, and the wave-mechanical models of the atom were proposed).
- The electron cloud describes the location of electrons in terms of probable regions with fuzzy boundaries (cloud).
  - Lower energy level electrons are closer to the nucleus while higher energy level electrons are further away
  - These levels need not be exact spherical regions, and thus the combination of many of these probable region forms the cloud.

Isotopes

- **Isotopes are atoms of the same elements with different masses**.
- All isotopes of a particular element have the same proton number, but different nucleon number (mass number) due to a different number of neutrons.
    - For example, the mass spectrometer shows that chlorine has 2 isotopes, chlorine-35 and chlorine-37.
- Isotopes have the same number of electrons and hence the same chemical properties (because chemical properties are fundamentally dependent on electron transfer and redistribution). However, isotopes only have different number of neutrons, thus:
    - They have different masses **(and thus different densities)**
    - They have different physical properties (boiling point, melting point)
    - For example, pure chlorine-37 has a higher mass, density, boiling and melting point than pure chlorine-35
- Relative atomic mass of isotopes:
    - If a sample only contained chlorine-35, then its Ar would be 35.
    - If a sample only contained chlorine-37, then its Ar would be 37.
    - If a sample contained chlorine-35 and chlorine-37 in the ratio 3:1, then its Ar would be 35.5 (as found in the periodic table).
    - Take the Ar of an element by taking the percentages of Ar against their abundance by percentage (refer to above example).
    - This explains why naturally occurring elements have relative atomic masses which are not close to a whole number, because they have naturally occurring isotopes, and their Ar would thus be a mean value taking into account the different isotopes and their relative proportion.
- Isotopes are either radioactive (radioisotopes) or non-radioactive.
- Isotopes are discovered by the mass spectrometer
    - The spectrometer has 3 basic components:
        - Ionizer: The gaseous sample is bombarded with electrons to create a positive ion of the element (charge of 1+)
        - Mass analyser: Ions of the same charge are separated according to their mass to charge ratio (m/Z) by using a magnetic or an electric field.
        - Data Analysis: The data is presented as a mass spectrum in which the m/Z is plotted against the % abundance. (The mass can be obtained from the m/Z as the Z, charge is simply +1)

Ionization Energy

Ionization Energy: The amount of energy required to convert 1 mole (6×10^{23}) of gaseous atoms to from 1 mole of positively charged (1+) ions.

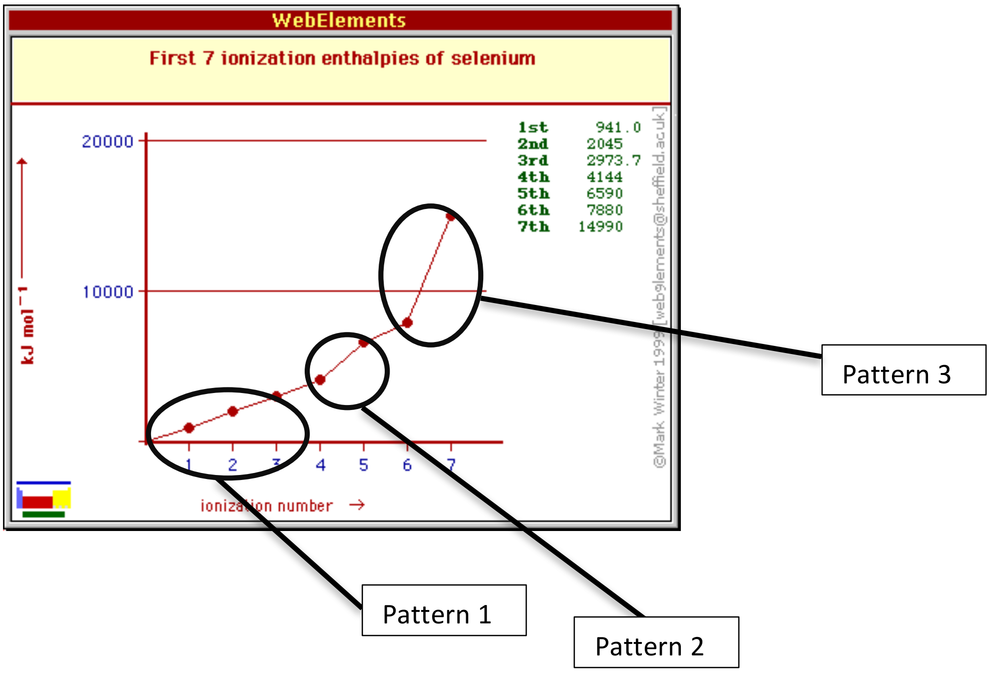
- Scientists make observations about electronic structure of atoms by observing the ionization energy.
- The first IE is the energy required to remove 1st electron from 1 mol of atoms, the second IE is the energy required to remove the 2^{nd} electron from 1 mol of the now positive ion, and so on.

- As successive electrons are removed, the ionization energy increases:
    - Big jumps in increase of ionization energy increase seen when all the electrons from one shell have been removed, and an electron from the next subshell is being removed.
    - Modest jumps of ionization energy increase seen when all the electrons from one subshell of an electron shell have been removed, and an electron from the next subshell is being removed.

<u>Patterns to note:</u>

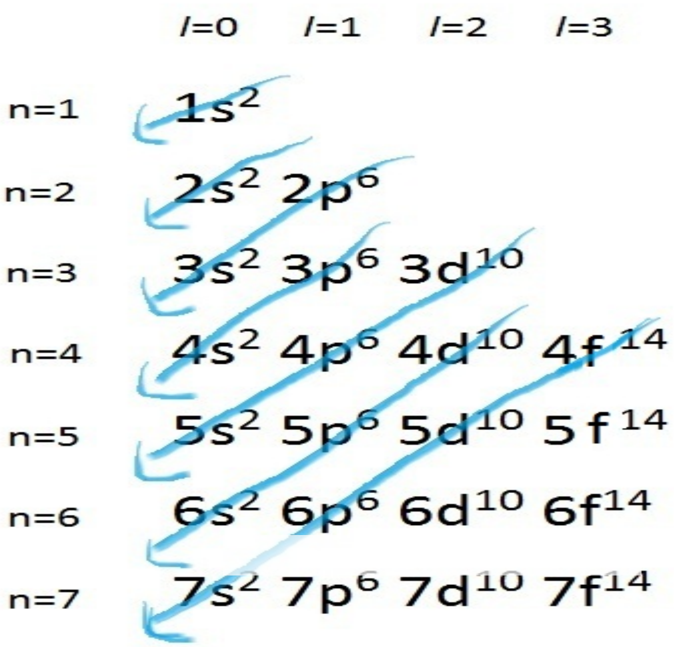
1. The ionization energies increase for each successive loss of electrons
2. There is a significant jump in the ionization energy whenever an electron is removed from a shell closer to the nucleus than the previous electron removed
3. There is a modest jump in the ionization energy whenever an electron is removed from a subshell closer to the nucleus than the previous electron removed

Graph of ionization energies of selenium:



Subshells

1. The shell number is known as the principal quantum number, and this describes the average distance of the electron from the nucleus.
2. Each shell consists of subshells labelled s, p, d, f. The number of subshells in each shell equals the shell number.
    a. The first shell has 1 subshell, s.
    b. The second shell has 2 subshells, s and p.
    c. Etc.
3. Each subshell has a number of orbitals – an orbital is the space in which the electron is most probably located.
    a. Subshell s has 1 orbital
    b. Subshell p has 3 orbitals
    c. Subshell d has 5 orbitals
    d. Subshell f has 7 orbitals
4. Each orbital can be filled by a maximum of 2 electrons, each with an opposite spin.
    a. But within an electron subshell (for example the P subshell which has 3 orbitals, 3 electrons would each occupy 1 orbital of their own as electrons would stay as far apart from each other as possible)
5. Electrons fill the subshell with the lowest energy first.
    a. Memorize the energy level of the subshells with this:



- When writing the electrons in a subshell, it is written as Mn^x where:
    - M is the shell number
    - n is the subshell alphabet
    - x is the number of electrons in that subshell
- When presenting answer, must present in order of principle quantum number (cannot write $4s^2$ in front of $3d^{10}$).

  - Write down the electronic configuration based on how it would be filled up
  - Re arrange it to present it in order of the principle quantum number

Molecular Mass

The relative atomic mass or relative molecular mass (Ar or Mr) is the mass of the atom/molecule over 1/12 of the mass of the C-12 atom (which is the same as the mass of a hydrogen atom. Note that the Ar and Mr do not have units.

Topic 4: Chemical Bonding and Structure

There are 3 types of strong chemical bonds – ionic bonds, covalent bonds, and metallic bonds. There are 3 types of weak bonds – instantaneous dipole-induced dipole interaction, permanent dipole-induced dipole interaction, and hydrogen bonding.

Note: **When asked to show all electrons, draw circles to denote each shell.**

Octet Rule

- Atoms achieve the configuration of an unreactive noble gas atom to give the compound stability.
- A minimal amount of energy is used to form such compounds (thus conferring stability).

2 exception:

- The first few elements of the periodic table will form the configuration of helium (which has 2 valence electrons in the first shell, called the duplet state). This is also the configuration of an unreactive noble gas.
- "Octet Expanders": Atoms in the 3rd period or higher can accommodate more than 8 electrons in some of their valence shells.

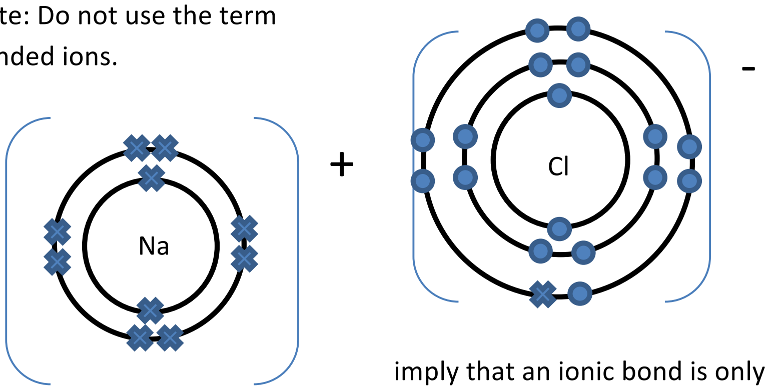
Note: Do not say "fully filled valence shell" to refer to the octet structure, because some shells such as shell 3 can have up to 18 shell, but are still unreactive when there are only 8 electrons (thus 8 electrons is not a fully filled valence shell).

4.1: Ionic Bonding

- Occurs between metallic and non-metallic elements
- Metal atoms (with 1, 2, or 3 valence electrons) lose their valence electrons to form positively charged ions (cations)
- Non-metal atoms (with 5, 6, or 7 valence electrons) gain electrons to form negatively charged ions (anions)
- Note: An atom with 1 valence electron will not gain 7 valence electrons, as it does not have enough protons to balance out the charge and hold in 7 extra valence electrons.
- Oppositely charged ions attract each other and thus the two ions are held together by an **electrostatic force of attraction**.

Note: A specific sodium atom does not transfer its electron to a specific chlorine atom. An ionic bond is formed between any oppositely-charged ions. In an ionic solid, each positively-charged ion is attracted to several negatively-charged ions and vice versa.

Note: Do not use the term *molecule* to refer to ionic bonded ions.

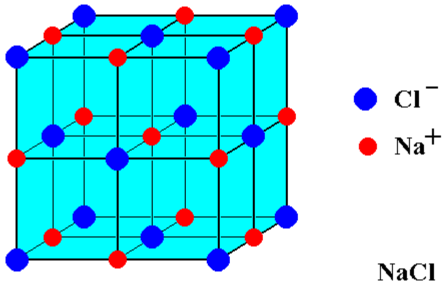


A dot-and-cross diagram is used to represent the electron arrangements within ions and to show the ratio of cations to anions. The diagram, however, does NOT imply that an ionic bond is only formed between the ions shown.

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.

4.2A: Covalent Bonding

- Occurs between non-metal atoms
  - Even with atoms of the same element
- Molecules can have single, double or triple bonds, in which each bond is a pair of 3 shared electrons
- Macromolecules occur when a large number of atoms are joined together by covalent bonds (diamond from carbon)

Simple Molecular Structures

- For molecules with more than two atoms, there tend to be one or two central atoms. **Always choose the element that has the least number of atoms in the formula to be the central atom.**
- Note that covalent bonding only exists within the structure, and there are weak intermolecular forces of attraction between molecules (Van der Waals' forces).
    - Thus, most simple molecular structures are liquids or gases at room temperature, and have low melting and boiling points, and do not conduct electricity in any state.

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.

How do covalent bonds hold the atoms together:

- In a hydrogen molecule, there are 2 protons, and 2 electrons. Each proton is attracted to each electron, and each electron is attracted to each proton (thus forming 8 forces of attraction), and each proton is repelled by the other proton (forming 2 forces of repulsion), and each electron is repelled by the other electron (forming 2 more forces of repulsion). Thus, the overall force is of attraction.

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.**

Essentially, the weak intermolecular forces of attraction between different molecules are not covalent bonds (and only those are broken during melting and boiling), but rather, covalent bonds are the bonds between the covalently bonded atoms, and these bonds are very strong.

Macromolecular Structure

Diamond and graphite are both macromolecules (millions of carbon atoms are bonded strongly together in a giant covalent structure). They are also allotropes (different forms of the same element) of one another. These allotropes exist because of the way the atoms are arranged and bonded.

- Diamond is the hardest natural substance know, whereas graphite is soft and is used as a solid lubricant to reduce friction in engines.
    - This is because diamond's structure consists of millions of carbon atoms strongly bonded in a tetrahedral (pyramid) structure, making diamond very hard as a substance.

  - Graphite's structure consists of parallel layers of hexagonal rings of carbon atoms. Although the carbon atoms are bonded very strongly with one another, the weak Van der Waals' forces between the layers can be easily broken, and allow the layers to slide against one another, making graphite soft and slippery.
- Diamond does not conduct electricity, while graphite conducts electricity.
  - In diamond, each of the carbon's 4 valence electrons is covalently bonded with other carbon atoms, thus there are no delocalized electrons to move through the structure to conduct electricity.
  - In graphite, each carbon atom is covalently bonded to three other carbon atoms, thus leaving one valence electron per carbon atom not involved in bonding. This electron becomes a free floating delocalized electron that can move freely along the layers of carbon atoms conducting electricity.

Whereas simple molecular structures have low melting and boiling points, macromolecules have different properties:

- High melting and boiling points
- They do not dissolve in any solvents
- The strong covalent bonding extends throughout the space of the molecule such that the whole crystal is considered to be one large molecule.

Other macromolecules:

- Si (silicon): Exists in a tetrahedral structure (exactly the same as diamond but 1 more shell)
- SiO_2 (sand): A giant covalent structure in a tetrahedral arrangement.
  - Note that each silicon atom is bonded to 2 oxygen atoms (that's why SiO_2).

4.2B Coordinate (Dative) Bond

A coordinate/dative bond is a covalent bond in which the shared pair of electrons is provided by only one of the bonded atoms (the donor), while the other atom (the receiver) does not contribute any electrons to the bond.

Note: **Once formed, a coordinate bond has the same characteristics as a covalent bond.**

Note: **For an atom to act as a donor, it must have at least 1 lone pair of unshared electrons in the valence shell. The acceptor must have at least 1 empty orbital (space for one electron pair) in its outer shell**.

Note: **When a positively charged ion forms a coordinate bond with a neutral covalent molecule, the charge applies to the whole ion, and thus the whole ion would have a +1 charge.**

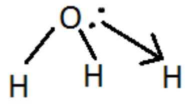
For example, in water, the oxygen atom within the H_2O molecule has 2 unshared pairs of electrons. This can be shared with an atom which needs them to complete its valence shell. A proton, H^+, has one empty 1s orbital. Hence, the H^+ accepts a pair of electrons from oxygen, thus achieving a full first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

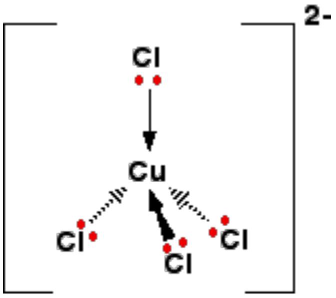
Dot and Cross Diagram:

- Instead of indicating a pair of shared electrons with a line, draw an arrow **from the donor to the receiver atom.**



Simple VS Complex Ions:

- Simple ions are formed when an atom gets a positive/negative charge.
- Complex ions are formed when a molecule gets a positive/negative charge.
- For example, coordinate bonds form between $4Cl^-$ and Cu^{2+} ions.



4.3A: Electronegativity

Ionic and covalent bonding are two extreme models of the chemical bond. Ionic compounds will still have some form of covalent characteristics, and covalent molecules will have some form of charge.

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. For example, when O and H form a covalent bond, O pulls the shared electrons a little towards itself (because O has more protons). This leaves the H with a partial positive charge, and the O with a partial negative charge, resulting in a **polar** covalent bond. It is said that oxygen is more electronegative than hydrogen.

Positive and negatively charged covalent substances are represented using $\delta + and\ \delta -$

- Greatest electronegativity: F (4.0), O (3.5), N(3.0)
- The rest will be provided during exams.

Case 1: In a covalent bond between two of the same atoms (Cl_2 or H_2 for example), both atoms are equally electronegative (both pull the shared electron pair equally). [Non-polar covalent]

Case 2: However, in a covalent bond between two atoms of different electronegativities, the atom with higher electronegativity will pull the shared electron pair slightly towards itself. [Polar covalent]

Case 3: If one atom is a lot more electronegative than the other, it will pull the shared electron pair to itself, and the other atom would have lost all control over that electron pair. [Ionic]

To determine which case a covalent substance falls into, the following rule of thumb is applied:

0 < Difference in electronegativity < 0.5	Non-polar covalent
0.5 < Difference in electronegativity < 2.0	Polar covalent
2.0 < Difference in electronegativity < 4.0	Ionic

Note: When comparing electronegativities between atoms, make use of the number of protons and the number of electron shells to support the argument.

4.3B: Polarizability

In ionic bonds, the protons of the positive ion are attract the electrons of the negative ion, thus the electron cloud of the negative ion becomes distorted in the direction of the positive ion.

- A negative ion that readily undergoes a large distortion is said to be highly polarizable
- A positive ion that can cause a large distortion is said to have high polarizing power.
- A larger negative ion (more electrons) is said to be more highly polarizable.
- A positive ion that is highly charged has a stronger polarizing power.

Case study: Although aluminium chloride is a compound formed between a metal and a non metal (thus, by right ionic), the strong polarizing power of Al^{3+} causes a pair of electrons from Cl^- to move so close to Al^{3+} that the bond is mainly covalent.

Note: The easier an atom can be polarized, the stronger the intermolecular forces that will form (thus, the higher the melting and boiling points).

4.4A: Dipole

Dipole**: A dipole is a separation of opposite charges.** Thus, in a polar bond that connects one slightly positively charged atom to another slightly negatively charged atom, a dipole is present (2 dipoles are present in a carbon dioxide molecule).

A dipole movement is the direction that of the negative partial charge.

- If the dipole movements are in opposite directions (such as in carbon dioxide), they cancel each other out and result in a non-polar molecule.
- If the dipole movements are at an angle to each other (such as in water molecules), they do not completely cancel, resulting in a polar molecule.
- In general, symmetrical molecules (like carbon dioxide) result in non-polar molecules, while asymmetrical molecules (like water molecules) result in polar molecules.
- A dipole movement is represented by an arrow with a "+" at its tail, in the direction of the negative partial charge.

4.4B: Instantaneous Dipole-Induced Dipole Interaction

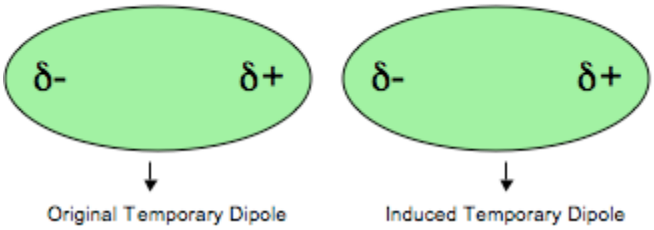
- Also known as Van der Waals forces – and is just a form of intermolecular force (only applies for covalently bonded compounds)
    - We say "weak intermolecular forces of attraction" for covalent compounds. This is one of them.
- Electrons in a molecule are in continual motion, and at any particular moment, the electron charge cloud around the molecule will not be perfectly symmetrical (there will be more negative charge on one side of the molecule than the other).
- This results in the occurrence of an instantaneous dipole.
- This instantaneous dipole induces a dipole formation in neighboring molecules.
- Thus, electrostatic attraction would exist between the instantaneous dipole and the induced dipole.

- The instantaneous dipole-induced dipole interactions are short lived because the electrons keep moving, so the dipoles vanish and reform.

The overall attraction in id-id interactions is weak (the weakest amongst the three intermolecular forces).

Factors affecting strength of id-id interactions:

- The greater the number of electrons in the molecules, the greater the strength of the id-id interaction.
- The greater the surface area of the molecules, the greater the strength of the id-id interaction



4.4C: Permanent dipole-permanent dipole Interaction

Polar molecules have permanent dipoles (the electrons don't keep shifting and re-arranging). Thus polar molecules re-arrange themselves such that the partially positive end of one molecule is near the partially negative end of another molecule. This is known as pd-pd interaction.

Example: Hcl molecules (H-Cl-------H-Cl--------H-Cl)

The electrostatic forces of attraction between the partially positive end of one molecule and the partially negative end of another molecule give rise to pd-pd interactions. Since these dipoles are permanent, pd-pd interactions are generally stronger forces than id-id interactions.

Note: Substances made up of polar molecules are held together by both pd-pd interactions and id-id interactions. Thus, they have higher melting and boiling points than substances made up of non-polar molecules.

Note: The strength of pd-pd interactions increases with an greater dipole movements (provided other factors such as number of electrons remains similar).

4.4D: Hydrogen Bonding

For hydrogen bonding to occur between 2 molecules, one molecule must have a hydrogen atom bonded to a F, O, or N, and the other molecule must have a F, O or N.

The hydrogen bond is an electrostatic force of attraction between the partially positive H and the lone pair of electrons on the F, O, or N.

Hydrogen has no inner core electrons (and only 1 proton), so when bonded to a highly electronegative F, O, or N, the shared electrons are drawn towards the F, O, or N. This causes the proton of the hydrogen to be partially exposed, which then interacts strongly with the lone pair of electrons of an electronegative atom (F, O, or N) of the other molecule.

Note: They atoms must be F, O, or N because these are the most highly electronegative atoms, and high electronegativity is required for hydrogen bonding (to sufficiently expose the hydrogen proton). Also, the lone pairs of electrons are not too spread out such that they diffuse in space.

Case studies:

- Hydrogen bonding may occur between the same type of molecules (hydrogen fluoride) or between different types of molecules (between hydrogen fluoride and $2H_2O$).

Note: Hydrogen bonding occurs between molecules containing hydrogen covalently bonded to a highly electronegative F, O, or N. These molecules also have pd-pd interactions and id-id interactions.

- Water has such an unusually high boiling and melting point because of the hydrogen bonding present.
- While most liquids become denser when they change to their solid forms, water becomes less dense because of the highly ordered 3 dimensional structure in ice which arises from hydrogen bonding. This arrangement of water molecules in ice creates a very open structure and prevents the molecules from getting too close to one another (thus making it less dense).
- Water freezes from top to bottom (snow-fishing): This is extremely helpful as it allows living organisms in a body of water to survive even in freezing conditions.

Note: A dotted line between the two lone electrons of the F, O, or N and the floating H is used to represent the hydrogen bond.

Note: In ionic compounds, the ionic bonds are broken during state change. In covalent substances, the abovementioned intermolecular bonds are broken.

4.5: Metallic bonding

In metal atoms, the valence electrons are shared equally by all the metal atoms, rather than being associated with a particular atom. These valence electrons break free from the atom, forming a sea of delocalized electrons, and making the metal atom a positively charged ion. The metal is held together by electrostatic forces of attraction between the positively charged ions and the negatively charged sea of delocalized electrons.

- This sea of delocalized electrons conducts heat and electricity throughout the metal.
- **Metal can be beaten into sheets (malleable) and twirled into wires (ductile) because the layers of atoms in the metal lattice structure can easily slide over each other**.
- Metallic bonds are strong – thus metals have high melting and boiling points.

Note: Metallic bonding forms giant structures.

Summary

- There are three types of giant structures: metallic structures held together by **metallic bonding**, macromolecular structures held together by **covalent bonding**, and ionic structures (as the ions are held in a giant lattice) held together by **ionic bonding**.
- We need to explain 3 properties of substances based on their structure and the type of bonding involved:
    - High or low melting point
    - Electrical conductivity
    - Solubility

	Ionic Structures	Simple molecular structures	Macromolecular Structures	Metallic Structures
Melting and boiling points	**High** because ions are held together by strong electrostatic forces of attraction between the positive and negative ions Large amounts of heat energy is needed to break these bonds	**Low** because the molecules are held together by weak intermolecular forces of attraction (id-id, pd-pd, or hydrogen bonds) These bonds can be broken with small amounts of energy	**High** because the strong covalent bonds (not inter molecular) between atoms in the lattice need to be broken to melt or boil the substance Large amounts of energy is required to break these strong covalent bonds	**High** because of strong electrostatic forces of attraction between the sea of delocalized electrons and the positive metal ions Large amounts of heat energy is needed to break these bonds.
Electrical conductivity	**Conduct electricity in the molten or aqueous states** because ions can move around freely in these states and act as charge carriers to conduct electricity. [Ions cannot do this in the solid state because the ions are in fixed positions]	**Do not conduct electricity in any state** because they are made up of neutral molecules (thus there are no charged particles that can carry charges and conduct electricity).	**Do not conduct electricity in any state** because macromolecules are made up of neutral atoms (no mobile charged particles that can carry charges and conduct electricity). **Exception: Graphite (4th electron of carbon is delocalized and this sea can act as a charge carrier to conduct electricity).**	**All metals conduct electricity in the solid and liquid states** because metallic lattices consist of positively charged ions surrounded by a sea of delocalized electrons which can carry charges and conduct electricity.
Solubility	**Most ionic compounds are soluble in water**	**Most covalent substances do not dissolve in**	**Not soluble in any solvent**	**Not soluble in any solvent** but many metals

	because water molecules are polar and can interact with the ions and cause them to separate (this dissolve). Non polar organic solvents cannot cause ionic compounds to dissolve	**water but dissolve in non-polar organic solvents** because the solvent and the solute can interact as both have weak intermolecular forces of attraction		react with water (very quickly). For example, sodium – this is not dissolving, but rather a rapid chemical reaction

Note: Lead(IV) chloride is an exception – it is a simple covalent substance, not an ionic compound.

Topic 5: Chemical Formulae and Equations

Chemical Formulae

Molecular Formula:

- The molecular formula shows all the atoms in one molecule.

Empirical Formula:

- The empirical formula shows the simplest ratio of the element present in a covalent substance (CH_2O for glucose)
- The empirical formula shows the ratio of the ions in ionic compounds (NaCl for sodium chloride)
- Note: Empirical formula is not used to refer to specific substances, as two different substances might have the same empirical formula ($C_6H_{12}O_6$ and $C_3H_6O_3$)

Structural Formula:

- The structural formula shows how the atoms are joined to one another, where each bond is represented by a straight line.

Common Chemical Formulae

- Diatomic Molecules (everything else that exists as a single element is monoatomic):

  hydrogen (H2)
  nitrogen (N2)
  oxygen (O2)
  fluorine (F2)
  chlorine (Cl2)
  bromine (Br2)
  iodine (I2)

- Acid Formulae (all aqueous):
  - Sulfuric Acid - H_2SO_4
  - Nitric Acid - HNO_3
  - Hydrochloric Acid - HCl
- Alkali Formulae (all aqueous):
  - Sodium Hydroxide – $NaOH$
  - Potassium Hydroxide – KOH
  - Aqueous Ammonia – NH_3

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	Ba^{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-}

Reactions:

- Acid + Metal → Salt + Hydrogen
- Acid + Base → Salt + Water
- Acid + Carbonate → Salt + Water + Carbon dioxide
- Alkali + Ammonium compounds → Salt + Ammonia Gas + Water

Naming of Salts:

- If hydrochloric acid is involved: _____________chloride
- If sulphuric acid is involved: ______________ sulphate
- If nitric acid is involved: ______________ nitrate
- If sulphurous acid is involved: ______________ sulphite

State Symbols:

Salt Type	Aq	S
Nitrate	All	
Chloride	All other chlorides	Lead (II), silver
Sulfate	All other sulfates	Lead (II), barium, calcium
Carbonates	Sodium, potassium, ammonium	All other carbonates
Hydroxides and Oxides	Sodium, potassium, ammonium	All other hydroxides and oxides

*Word equation is usually provided.

Topic 6: Mole Concepts

Definition

1 mol is 6×10^{23}.

Molar Mass

The molar mass is the mass of 1 mol of a substance (essentially the Mr or Ar). Mr and Ar are relative molecular/atomic mass, and they have no units. Molar mass is the number of grams per mol (gmol^{-1}) – it has a unit!

Empirical Formula from Percentage Mass

Step 1: Let the mass of the substance be 100g. Find out the mass of each element in this 100g sample (so if H is 2.4%, there will be 2.4g of H in 100g of sample).

Step 2: Determine the Ar of each of the elements (from the periodic table).

Step 3: Find the amount of each of the substances (in mol). Divide the total g in 100g sample by the molar mass to find the amount in mol.

Step 4: Divide each amount in mol by the smallest amount in mol to obtain a ratio of elements (rounding off may be requires*)

Step 5: Simplify this ratio and write out the empirical formula

***If a whole number is not attained after dividing, then the ratio must still be set to a whole number. For example, if 1.2:1 is obtained, then it must be written as 5:4. However, if the number is very close to a while number, like 1.05:1, then simply round it off to 1:1**

Note: If the actual mass is provided (instead of percentage by mass), then just write "Mass (in g)" instead of "Assuming 100g, mass".

Note: The molecular formula can be obtained by multiplying the empirical formula by the factor (Mr of molecule/Mr of empirical formula)

Percentage Composition by Mass

To find out the percentage composition of an element or a compound in another compound, **take the Mr/Ar of the desired compound/element, and divide it by the Mr of the whole substance.**

Example: Percentage composition by mass of hydrogen in water is $\frac{2\times1}{18}\times100\% = 11.1\%$.

Rate of Diffusion

Denser gases diffuse slower than less dense gases. **Therefore, gases with larger Mr diffuse slower than gases with smaller Mr**.

Out of sulfur dioxide, carbon monoxide, hydrogen, and ammonia: Sulfur dioxide diffuses the slowest, followed by carbon monoxide, ammonia, and hydrogen.

Limiting Reactant

The *excess reactant* is the reactant that does not get used up. The *limiting reactant* is the reactant that gets completely used up.

Step 1: Calculate the number of moles of each reactant based on their given masses (in grams), using the Ar/Mr of the reactants.

Step 2: Compare the ratio of moles of the reactant used to the theoretical ratio given in the equation to determine the limiting reactant and the excess reactant.

Molar Volume of Gases

Avogardo's Law: Equal volumes of all gases under room temperature (25^{o}C) and pressure have the same number of atoms and molecules

1 mol of any gas occupies 24dm^3. This volume is thus known as the molar volume. Note that 1dm^3 is 1l (litre), and 1000cm^3.

Thus, the ratio of volumes reacted/produced in a reaction would be equal to the ratio of the amount of reactants/products.

Calculations Involving Masses

Step 1: Convert all reactants/products from given mass (in grams) to amount (in mol).

Step 2: Read off the theoretical proportion from the equation.

Step 3: Apply this theoretical proportion to the given amount to find the result (in mol).

Step 4: Convert the result back into grams using the Mr.

Percentage Purity

Percentage Purity = Mass of pure substance in sample/Mass of total sample X 100%

Percentage Yield

Percentage Yield = Actual Yield (in mol)/Expected Yield (in mol) X 100%

Concentration of Solution

Molarity of a solution is the concentration of the solute in the solution (in moles). Therefore, molarity (in mol/dm^3) is given by (Number of moles of solute (in mol)/Volume of solution (in dm^3))

The concentration of a solution by mass (in g/dm^3) is given by: (Mass of solute (in g)/Volume of solution(in dm^3)).

Conversion: Concentration of a solution by mass = Molarity of solution (mol/dm^3) X Molar mass of solute (g/mol)

Topic 7: Acids and Bases

General Notes

- When acids and bases are dissolved in nonpolar solvents like toluene, they do not exhibit acidic/alkaline properties. Needs to be dissolved in water

Chemical Equations (Acids)

1. **Acid + Base → Salt + Water**
    a. Hydrochloric acid + potassium hydroxide → potassium chloride + water
    b. Sulfuric acid + sodium hydroxide → sodium sulphate + water
    c. Sulfuric acid + potassium hydroxide → potassium sulfate + water
    d. Nitric acid + sodium hydroxide → sodium nitrate + water
    e. Nitric acid + calcium hydroxide → calcium nitrate + water
2. **Acid + Carbonate → Salt + Water + Carbon Dioxide**
    a. Hydrochloric acid + sodium carbonate → sodium chloride + water + carbon dioxide
    b. Hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide
    c. Sulfuric acid + magnesium carbonate → magnesium sulfate + water + carbon dioxide
    d. Nitric acid + iron (II) carbonate → iron (II) nitrate + water + carbon dioxide
3. **Acid + Metals → Salt + Hydrogen**
    a. Hydrochloric acid + magnesium → magnesium chloride + hydrogen
    b. Sulfuric acid + iron → iron sulfate + hydrogen
    c. Nitric acid + zinc → zinc nitrate + hydrogen
4. **Acid + Sulfites → Salt + Water + Sulfur Dioxide**
    a. Hydrochloric acid + sodium sulphite → sodium chloride + sulfur dioxide + water
    b. Sulfuric acid + potassium sulphite → potassium sulfate + sulfur dioxide + water
    c. Nitric acid + sodium sulphite → sodium nitrate + sulfur dioxide + water

Chemical Equations (Bases)

Note: Metal oxides and hydroxides are bases

1. **Acid + Base → Salt + Water**

    a. Refer to previous examples
2. **Ammonium Salt + Base → Salt + Water + Ammonia Gas**
    a. Sodium hydroxide + ammonium chloride → Sodium chloride + water + ammonia gas
    b. Calcium hydroxide + ammonium nitrate → Calcium nitrate + water + ammonia gas
    c. Magnesium oxide + ammonium sulfate → Magnesium sulfate +water + ammonia gas
    d. Iron (II) oxide + ammonium chloride → Iron chloride + water + ammonia gas
3. **Soluble base + Salt → Salt + Insoluble Hydroxide**
    a. Sodium hydroxide + copper (II) sulfate → sodium sulfate + copper (II) hydroxide
    b. Potassium hydroxide + aluminium sulfate → potassium sulfate + aluminium hydroxide

Strong/Weak Acids & Bases

Strong Acid	- Hydrochloric acid (HCl) - Nitric acid (HNO_3) - Sulfuric acid (H_2SO_4)
Weak Acid	- Ethanoic/acetic acid (CH_3COOH) - Lactic acid ($CH_3CHOHCOOH$) - Carbonic acid (H_2CO_3)
Strong Alkali	- Sodium hydroxide (NaOH) - Potassium hydroxide (KOH) - Calcium hydroxide ($Ca(OH)_2$)
Weak Alkali	- Ammonia (NH_3)

Note: Ammonia is NH_3 while ammonium is NH_4

Classification of Acids

Acids dissolve in water, and the acid molecules become ionized.

For acid 'HA', the general formula is: HA (liquid) + H_2O (liquid) → H_3O^+ (aqueous) and A^- (aqueous)

- In hydrochloric acid: H_3O^+ (aqueous) and Cl^- (aqueous)
- In ethanoic acid: H_3O^+ (aqueous) and CH_3COO^- (aqueous)

Rule of thumb (strength):

- **A strong acid has acid molecules that are completely ionized in aqueous solution**
- **A weak acid has acid molecules that are partially ionized in aqueous solution**
    - A weak acid has high proportions of the acid undissociated in water
    - To represent an equation of a weak acid, draw a double-faced arrow. This is because they can easily undergo a backward reaction to form the original acid.
    - For example: $HA + H_2O \leftrightarrow H_3O^+$ and A^-

Rule of thumb (concentration):

- The concentration of an acid refers to the amount of acid molecules dissolved in the solvent.

  - An acid of 0.100 $moldm^{-3}$ is considered dilute in comparison to an acid of 2.00 $moldm^{-3}$ which is considered concentrated.

Rules of thumb (basicity):

- Acids are monobasic (monoprotic), dibasic (diprotic), or tribasic (triprotic) – depending on the number of hydrogen ions that are able to ionize to form H^+ ions
  - HCl is monobasic
  - H_2SO_4 is dibasic
  - H_3PO_4 is tribasic
- Note that not all hydrogen atoms will ionize to form H^+ ions. Especially in organic acids, where usually only the last atom will ionize. Examples are ethanoic acid (CH_3COOH), and lactic acid ($CH_3CHOHCOOH$). These acids are monobasic.
- The basicity of an acid does not determine its strength, as the number of hydrogen ions liberated per molecule of acid does not determine its strength.
  - For example, carbonic acid is a weaker dibasic acid than nitric and hydrochloric acids which are both monobasic
- When a base is reacted with a dibasic or tribasic acid if the charge of the metal ion is (+1), then a few salts can form, for example:
  - $AgX + H_2SO_3 \rightarrow AgHSO_3$ (acid salt) or AG_2SO_3 (normal salt)
  - Multiple reactions can occur in such an acid-base chemical reaction, for example:
    - $H_2C_2O_4 + KOH \leftrightarrow KHC_2O_4 + H_2O$
    - $KHC_2O_4 + KOH \leftrightarrow K_2C_2O_4 + H_2O$
    - Overall equation: $H_2C_2O_4 + 2KOH \leftrightarrow K_2C_2O_4 + 2H_2O$

Classification of Oxides

Metal Oxides		Non-metal Oxides	
Basic oxides	Amphoteric oxides	Acidic oxides	Neutral oxides
React with acids to form salts	React with acids and bases to form salts All amohoteric oxides are insoluble	React with bases to form salts	Do not react with acids or bases
Sodium Oxide, Na_2O Calcium Oxide, CaO Copper(II) Oxide, CuO	Zinc Oxide, ZnO Aluminium Oxide, Al_2O_3 Lead(II) Oxide, PbO Only these 3 in syllabus	Carbon Dioxide, CO_2 Sulfur Dioxide, SO_2 Nitrogen Dioxide, NO_2	Carbon Monoxide, CO Nitric Oxide, NO Water, H_2O

Note: All hydroxides behave exactly the same way as oxides. As such, all these oxides can be replaced by the respective hydroxides.

- All metal oxides/hydroxides except zinc, aluminium, and lead are basic – (remember 'ZAP)
- Some non-metal oxides are acidic

pH of a Solution

The pH of a solution is a quantitative measure of the extent of alkalinity or acidity of a solution.

As **pH = -log$_{10}$[H$^+$]**, the pH of a solution is a measure of a solution's concentration of hydrogen ions, and these values fall between 0 to 14.

- A change of 1pH changes the hydrogen ion concentration by a factor of ten. In other words, a solution of pH 1 has ten times the concentration of (and so, then times more) hydrogen ions than a solution of pH2.
- pH values can extend beyond the range, but not very often
- An aqueous solution always contains H^+ and OH^- ions, due to the small extent of partial dissociation of water molecules
  - $H_2O \leftrightarrow H^+ + OH^-$
- Misconception: Many students think that acidic solutions do not contain OH^- ions. This is not true. In an acidic solution, the concentration of H^+ ions is greater than the concentration of OH^-, and thus it is acidic.

Measuring pH values:

- Chemical indicators
  - Indicators that show different colours at different pH values
  - Universal indicator, litmus paper
- pH meter/sensor
  - Electrical device that is more accurate than indicators – but requires calibration and is dipped into the solution directly

pOH: A measure of the concentration of the hydroxide ions. In essence, **pOH = 14 – pH**

Models of Acids and Bases

Arrhenius' Model

Theory: Acids are compounds that ionize when they dissolve in water to give H^+ ions and a corresponding negative ion. Bases are compounds that ionize in water to give OH^- ions and a positive ion.

- Arrhenius acid: A substance that ionizes when it dissolves in water to give a hydrogen ion
- Arrhenius base: A substance that ionizes when it dissolves in water to give a hydroxide ion

Strengths:

- It explains why all acids & bases have similar properties to one another (because they dissociate in water to form a certain ion)
- It explains the observation that acids and bases counteract each other (neutralization)

Weaknesses:

- It can only be applied to reactions in water – because it defines acids and bases as per what happens when the compounds dissolve in water
- Only compounds that contain the OH^- ion can be classified as Arrhenius bases, as his model fails to explain why compounds such as NH_3 have the characteristic properties of bases

- The existence of a H^+ ion is doubtful, for it is so small yet generates an electric field so intense that it attracts any molecule with any unshared pairs of electrons. For example, $H^+ + H_2O \rightarrow H_3O$

Bronsted-Lowry Model

Theory: While Arrhenius' general theory was accepted, it came to be known that a H+ ion (a proton) cannot and does not exist on its own for more than a tiny fraction of a second. Hence, no matter how strong the acid, the concentration of hydrogen ions in an aqueous solution would be so small as to be practically non-existent.

- Acid: A substance which can donate a hydrogen ion (a proton) to another substance
- Base: An substance that can accept a hydrogen ion from another substance

Strengths:

- The Bronsted-Lowry definition of acids encompasses all Arrhenius acids (since a substance that can ionize to form hydrogen ions can also donate hydrogen ions), but also extends to acid-base reactions that are not conducted in water, such as in gases.
- The Bronsted-Lowry definition of bases explains why substances that do not contain OH^- can act like bases, simply because bases are just opposite of acids (accept H^+)
  - By this definition, ammonia (NH3), baking soda, and insoluble metallic oxides (MgO) are considered bases, though by Arrenius' definition, they are not

Acid Rain

Acid rain refers to the deposition of acid in two parts, which falls back onto the Earth's surface in equal proportions:

- Wet: Wet deposition refers to acidic rain, fog, and snow, and this wet deposition has effects on all the varieties of plants and animals it flows over.
- Dry: Dry deposition refers to acidic gases and particles. When this dry deposition comes into contact with water/other forms of surface runoff, the combination of the depositions becomes stronger than the wet deposition alone

Normal rain is weakly acidic due to the presence of slightly acidic gases in the air. For rain to be considered acid rain, its pH needs to be **lower than 4**.

Steps involving the production of acid rain:

1. Source of pollutants
   a. Sulfur dioxide is produced by the combustion of sulfur containing fossil fuels in power stations and motor vehicles.
   b. Oxides of nitrogen are produced in car engines
2. Effects of the pollutants
   a. Sulfur dioxide and oxides of nitrogen damage the lungs and cause bronchitis
   b. In the atmosphere, the pollutants combine with water to form strong acids, such as sulphuric acid (from sulfur dioxide) and nitric acid (from oxides of nitrogen)

    c. Acid rain of these strong acids damages buildings, and lowers the pH of water bodies, thus killing aquatic life

3. Minimizing the effect:
    a. Sulfur dioxide pollution can be minimized by removing sulfur from fossil fuels before burning them, or by removing sulfur dioxide from the waste gases produced during fossil fuel combustion
    b. Oxides of nitrogen pollution can be minimized by fitting cars with catalytic converters.
4. Removing acids from environments:
    a. Adding lime (CaO) or slaked lime ($Ca(OH)_2$) to soil/water bodies neutralizes the acid, but goes on to render the environments alkaline
    b. Adding limestone ($CaCO_3$) neutralizes the acid, but since it is insoluble in water, remains in the solid form and does not render the environments alkaline

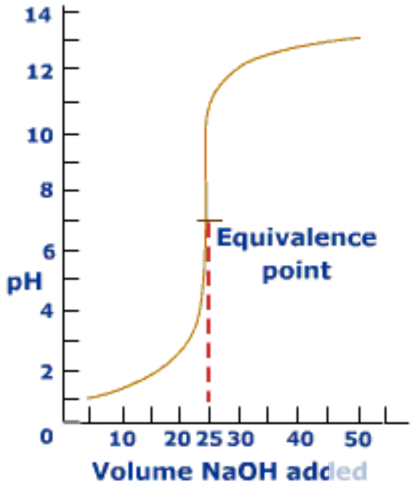
Acid-Base Titration Curves

Rules of thumb:

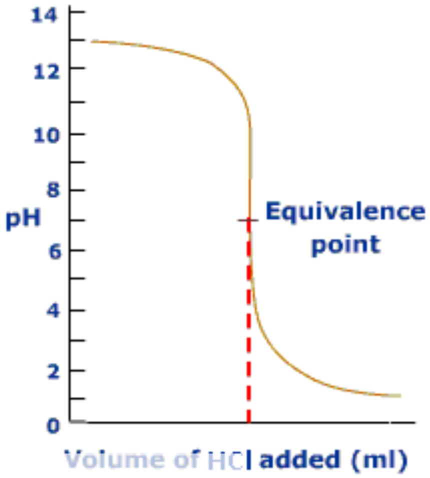
- Strong acid (pH 0 to 1) + strong base (pH 13 to 14) = pH 7 solution
- Strong acid (pH 0 to 1) + weak base (11 to 12) = pH 5 solution
- Weak acid (2 to 3) + strong base (pH 13 to 14) = pH 9 solution

Note: If it starts with acid, use the acidic end of the alkali scale to end with. If it starts with alkali, use the alkaline scale of the acid to end with.

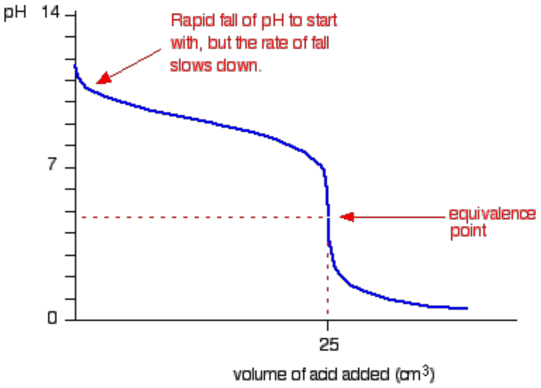
Note: Whichever one is weaker, that one will have the less steep gradient.

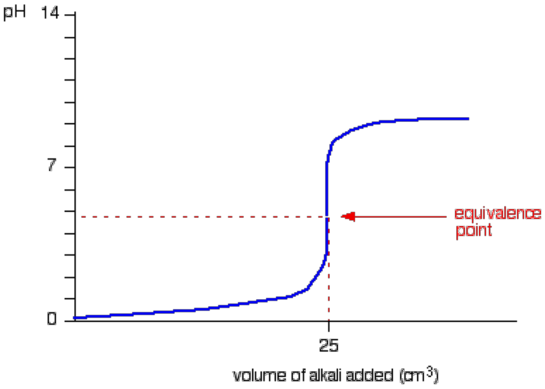


Titration curve of strong acid (HCl) with a strong base (NaOH)



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





Applications of Neutralization

- Controlling pH of soil – to reduce acidity of soil, farmers add lime/slaked lime to soil.
- Treatment of indigestion – to relieve the pain of too much hydrochloric acid, a mild alkali such as magnesium hydroxide is taken
- Treatment of insect stings – acid from bee stings is neutralized with baking soda
- Treatment of industrial waste – acidic waste is treated with limestone

Types of Observations (Acids and Bases)

1) Production of a gas	"Effervescence is observed"
2) Production of a solid	"A ____ precipitate is formed"
3) Dissolving of a solid to form a solution	"The ____ solid dissolved to form a _____ solution"
4) Colour change of the pH indicator	"The ____ litmus paper turned ____"

Solubility Rules

1) **Chlorides, bromides, and iodides are soluble in water, except lead and silver.**
2) **Most sulfates are soluble, except barium, lead (II), and calcium.**
3) All carbonates are insoluble except ammonium, and group I carbonates (hydrogen, lithium, sodium, potassium, etc)
4) All oxides and hydroxides are insoluble except group I oxides/hydroxides (hydrogen, lithium sodium, potassium, etc) and calcium.
5) All ammonium compounds and group I compounds (hydrogen, lithium, sodium, potassium, etc) are soluble.
6) All nitrates are soluble.

Ammonium compounds / Group I compounds / Nitrates – Soluble.

All other oxides, hydroxides, & carbonates – Insoluble.

Steps:

1. Check if substance is a nitrate, a group 1 compound, or an ammonium compound. If it is, then it is soluble
2. Check chloride and sulfate exceptions. If it is one of the exceptions, insoluble.

3. Else:
    a. If chloride/bromide/iodide/sulfate – then soluble
    b. If carbonate/oxide/hydroxide – then insoluble

Precipitation Reaction

A precipitate is the formation of an insoluble solid when 2 aqueous solutions are mixed.

Requirements:

- One reactant must contain the positive ion of the precipitate, and one reactant must contain the negative ion
- Both ions need to be in the aqueous states
- Note: A precipitate should not be formed by reacting a solid with a solution

Collection of precipitate:

1. Filter the mixture
2. Wash residue on filter paper with distilled water
3. Dry the solid by pressing it between sheets of filter paper

Ionic Equations

When substances react, not all the atoms/ions/molecules take part in the reaction. As such, an ionic equation specifically shows which atoms/ions/molecules take part in a reaction. The aim of an ionic equation is to eliminate the spectator ions (free ions present at the start of the reaction, and remain unchanged by the end of it).

Steps:

1. Write out the balanced chemical formula with state symbols (use solubility rules)
2. Write out the free ions in the equation
    a. Note that substances in solid, liquid, or gaseous states do not form free ions
    b. State symbol of the free ion is the same as that of the substance it originates from (aq)
3. Cancel out the spectator ions (ions that remain unchanged at the end of the reaction)

Note that the number of each type of atoms, as well as the net charge on both sides of the equation must be balanced.

- In general, all the ionic equations for neutralization reactions (acid-alkali reactions) are the same: $H^+ (aq) + OH^- (aq) \rightarrow H_2O (l)$
- In general, the ionic equation of a precipitation reaction involves only the ions of the precipitate.
- In general, the ionic equation of an acid + metal reaction involves only the ions of the metal and the hydrogen.
- In general, the ionic equation of a carbonate reaction only involves the hydrogen and carbonate ions.

- In general, the ionic equation of a displacement reaction only involves the ions that are reduced and oxidized.

Redox Reactions

A redox reaction is one in which oxidation and reduction occur simultaneously in a complementary manner; one substance is reduced and one substance is oxidized. The reduced substance is the oxidizing agent and the oxidized substance is the reducing agent.

- A substance is oxidized when its oxidation number increases (when it loses electrons)
- A substance is reduced when its oxidation number decreases (when it gains electrons)
- Remember: OIL RIG (oxidation is loss, reduction is gain)

Half Equations

- Each redox reaction can be seen as two parts (oxidized and reduced), each with a half equation.
- Take the substance of concern, and examine it before and after the reaction.
- Add the relevant electron(s) at the appropriate location

Oxidation Number Rules

1. Elements that are uncombined with other elements are assigned an oxidation number of 0.
2. In simple ions, the oxidation number is simply the charge of the ion.
3. In polyatomic ions, the sum of oxidation number is equal to the net charge.
4. The oxidation number of hydrogen in all its compounds is +1, except in metal hydrides where its oxidation number is -1.
5. The oxidation number of oxygen in all its compounds is -2, except in peroxides where its oxidation number is -1.
6. The oxidation number of Group I elements (such as sodium) in their compounds is +1, for Group II elements (such as magnesium) in their compounds is +2, and for aluminium in its compounds is +3
7. There are many oxidation numbers for Group VII elements in their compounds but the usual one is -1.
8. The sum of the oxidation numbers of all the elements in a molecule or a compound is 0.

Lab Tests for Oxidizing/Reducing Agents

Reducing Agents:

1. Add a few drops of acidified potassium manganite VII ($KMnO_4$) to unknown substance
    a. If purple reagent turns colourless: unknown substance has reducing agent

Oxidizing Agents:

1. Add a few drops of aqueous potassium iodide (KI) to unknown substance
    a. If colourless reagent turns brown, unknown substance has oxidizing agent

Balance a Redox Reaction using Half-Equations

1. Write the unbalanced equation in the ionic form, and derive the two half-equations from it.
2. Balance the atoms in the half equations:
    a. Use H_2O and H^+ to balance if in acidic solution
    b. Use H_2O and OH^- to balance if in alkaline solution
3. Add electrons to balance the charges
4. Multiply the equations such that the electrons added in both equations is the same.
5. Add the half equations together and subtract the terms that appear on both sides of the equation.
6. If spectator ions are known, add them into the final equation.