Chemistry Notes MYCT 2013

Nature and Classification of Matter

Types of particles: Atoms, Molecules and Ions (which usually exist in the plasma form)

Elements/Compounds/Mixtures:

- Element: A substance that is made up of only one type of atom. It **cannot be chemically split** into simpler substances. (Note: Radioactive decay may occur which splits atoms into smaller atoms, however nuclear means should not be confused with CHEMICAL means-which refers to heating and other chemical methods)
- Atom of an element: a particle representing an element that is NOT chemically bonded with other atoms (such as a lone helium or argon atom)
- Molecule of an element: particle representing an element that is **chemically bonded** with other atoms of the same element (eg. H2, O2)

- Compound: A substance that contains **two or more elements** chemically joined together (through reacting of substances) (such as water)
- Molecule of a compound: Particle representing a compound with 2 or more atoms chemically bonded together (the atoms must be of 2 or more different kinds)- eg. HCl, H2O)

- Mixture: A substance that contains two or more substances physically together but not reacted with one another chemically (such as oxygen and hydrogen together)
- **Mixture of elements**: Consists of 2 or more elements NOT chemically bonded together (eg. Hydrogen and Helium)-BRONZE and STEEL (alloys)
- **Mixture of Elements and Compounds**: Consists of 1 or more elements NOT chemically bonded with 1 or more compound (eg. Water and Oxygen: whereby water is a compound while oxygen is an element)- AIR
- **Mixture of Compounds**: 2 or more compounds NOT chemically bonded together (eg. Carbon Dioxide and Water)- PETROL

Note (Alloys):
  - Brass=Copper and Zinc
  - Bronze=Copper and Tin
  - Steel=Iron and Carbon
  - Solder=Lead and Tin

- Atom: basic matter of unit (dense central nucleus surrounded by a cloud of negatively-charged electrons)
- Molecule: Electrically neutral group of two or more atoms held together by chemical covalent bonds
- Ions: Electrically charged particles that are exist in metallic or ionic bonding

Elements vs Compounds

Similarities	Differences
Both are made up of atoms	An element is made up of only one type of atom (a compound is made up of more than one type of atom)
Both are pure substances (eg. Exists independently without interference from other substances-which forms a mixture)	An element can be metallic or non-metallic (compounds can be formed by combining a metallic element with non-metallic elements)
Both have fixed melting and boiling points	An element cannot be chemically split into simpler substances (compounds can be decomposed/electrolyzed into smaller compounds)

Misconceptions:
- Molecules are heavier than atoms (Untrue: An uranium atom is heavier, which means a greater mass, than an oxygen molecule)
- A molecule is equivalent to a compound (molecules are formed through covalent bonding while compounds can be formed through any bond)

Compounds and Mixtures

Similarities	Differences
Both are made up of more than one type of atom (even mixtures combine 2 or more types of atoms together: helium and hydrogen)	A mixture can be separated into its components through physical methods (A compound can only be separated through chemical methods)
	A mixture does not have fixed melting or boiling points-because of concentration of components (compounds have fixed melting and boiling points)
	No chemical changes take place when mixtures are formed (chemical changes occur when compounds are formed)
	Proportion of constituents in mixture can vary (elements are always combined in fixed proportion for compounds)
	Properties of mixtures are the same as those of its constituents (properties of compounds differ from constituents elements-affected through chemical changes)

Distinct Characteristics between Metals and Non-Metals

Metals	Non-metals
Ability to conduct electricity	Usually not able to conduct electricity (with the exception of graphite)
High melting and boiling points (with the exception of Mercury and Group I Alkali Metals)	Low melting and boiling points
Malleable and Ductile (due to metallic bonding structure-look for metallic bonding)	Not malleable or ductile
Group I to III/IV (with Polonium in VI and other exceptions)	Group IV to Group 0
Lustrous and Sonorous	Not lustrous and sonorous (generally)

Note: Be careful of Metalloids and Semi-Metals (like Germanium since they share properties from both metals and non-metals)

Particulate Nature of Matter:

- Main 3 states: Gas, Liquid, Solid
- 4th State of Matter: Plasma (exists at extremely high temperatures and in the form of ions, possibly through the ionization of gases)
- 5th State of Matter: Bose-Einstein Condensate (dilute gas of bosons cooled to low temperatures near absolute zero)

	Solid	Liquid	Gas
Arrangement of Particles	Orderly	Disorderly	Disorderly
Energy Content (derived from speed of particles)	Low	Moderate	High
Distance between Particles	Closely packed	Slight further apart	Very far apart (furthest)
Forces of attraction between Particles	Strong	Weak	Very weak
Movement of Particles	Vibration about a fixed point	Moving freely/randomly (vibration and translation)	Moving freely/randomly (vibration and translation)
Speed of Particles	Slowest	Faster than Solid Particles	Fastest
Compressible	No	Yes (only very slightly to remove the gaseous particles between molecules)	Yes

Phase Transition:

- Solid-Liquid (Melting and Freezing)
- Solid-Gas (Sublimation and Deposition)
- Liquid-Gas (Evaporation/Boiling and Condensation)
- Gas-Plasma (Ionization and Deionization)
- Particles gain energy to transition from solid to liquid, liquid to gas
- Particles lose energy to transition from liquid to solid, liquid to gas

Lower and Upper Fixed Points

- Pure substance is heated or cooled down, there will be 2 fixed points (boiling or melting points)
- Melting point: pure substance will change from solid to liquid state
- Boiling point: pure substance will change from liquid to gaseous state (the converse is true)
- Pure substances can only have 2 fixed points

Identifying impure substances:

- 3 or more fixed points
- Change of substance across of a range of temperatures
- Fixed points is different from usual
- Impurities usually raise the boiling point and lower the melting points of pure substances
- Usually many good effects in our daily lives (such as impurities added to car engine coolants)

Change of State:

- Liquid water upon gaining energy, uses it to increase the kinetic movement for particles
- At 100 degrees Celsius, the heat energy is supplied to break the intermolecular bonds between water molecules (to transition it to the gaseous state): thus, the temperature remains constant

Salt Water:

- Upper fixed limit increases (attraction between impurity and water: water does not boil easily since more heat energy is required to break the bonds)
- Lower fixed limit decreases (difficult for the orderly arrangement of water molecules: makes disorderly arrangement easier)

Diffusion:

- Refers to the **random movement** of particles from a region of **higher concentration** to a region of **lower concentration**
- Evidence of movement in liquid and gaseous states (not in solid state since the particles are closely packed together)
- Experiment 1: Use of bromine gas (brown and toxic: boiling of about 30 degree Celsius)
- Experiment 2: Use of potassium permanganate crystals (purplish-blue colour)
- In a closed system, there will be an **equal concentration** in the end (since the experiment does not interact with its environment)
- Happens very quickly in a vacuum (no obstruction from gaseous particles)
- Gravity does not aid diffusion (horizontal method is usually the most effective)
- Affected by temperature (higher temperature=faster diffusion)
- Greater mass and density of the molecule=slower diffusion
- Affected by humidity (higher humidity=slower diffusion)

Speed of diffusion:

- K constant X 1/ sq rt (Molecular/Atomic Mass)
- Speed of diffusion of Helium Gas/Speed of Diffusion of Oxygen Gas = 1/ sq root 4 / 1/ sq root 32 = 2.8 (times faster diffusion of helium compared to oxygen)
- Always compare Faster Gas/Slower Gas
- Unit is in m/s

Experimental Techniques and Designs:

Measurement Apparatus:
Burette:
- Maximum of 25 cm3
- Accurate to 2dp (second dp either 0 or 5)

Pipette:
- Usually measured up to 1dp
- Usually with a fixed volume (for the uptake of a solution)

Measuring Cylinder:
- Measured up to 1dp (with the dp being 0 or 5)
- Maximum of 100cm3

Beaker:
- Least accurate
- Usually 10-50 cm3 demarcations (which are also an estimate)

Downward Displacement of Water:
- For gases which are insoluble in water (such as nitrogen, hydrogen or methane)
- Water level goes down

Downward Delivery:
- For gases that are denser than air (carbon dioxide, sulfur hexafluoride, chloride, bromide)
- Rubber bung and stopper cannot be used (or the pressure will increase beyond the capacity)

Upward Delivery:
- For gases that are less dense than air (Helium/Hydrogen)
- Rubber bung and stopper cannot be used

Gas syringe:
- Collection of gas that is produced (for the testing of gases)
- Gas syringe must be graduated (to measure the exact volume of gas collected)

Drying of Gases:
- Calcium chloride (for all gases: acidic, alkaline, neutral)
- Anhydrous Copper II sulfate (removal of water: which turns the copper II sulfate into a blue colour)
- Calcium Oxide-an alkaline substance (not used for acidic gases)
- Concentrated Sulfuric Acid-an acidic substance (not for alkaline gases)
- Silica gel: for removal of water as well

Neutral Gases: Hydrogen and Oxygen
Acidic Gases: Sulfur Dioxide, Carbon Dioxide and Nitrogen Dioxide (all can become acids)
Alkaline Gases: Ammonia (which can become an alkali)

Identification of Gases:
Chlorine (Cl2):
- Use of manganese oxide in a dry test tube: add hydrochloric acid
- Pale green/yellow gas will evolve (chlorine)
- Blue litmus paper will turn red and then bleached (acidic)
- Chlorine is denser than air (downward delivery)

Hydrogen (H2):
- Add zinc and sulfuric acid (add copper sulfate as a catalyst)
- Insert a **lighted** splint into the test tube (gas extinguishes the splint with a ‘pop’ sound)
- Colourless gas evolved

Oxygen (O2):
- Add manganese oxide and hydrogen peroxide
- Insert **glowing** splint into the test tube
- Gas rekindled (relighted) the glowing splint

Sulfur Dioxide (SO2):
- Add sodium sulfite (Na2SO3) and hydrochloric acid
- Filter strip with potassium manganate at the mouth of the test tube
- Potassium manganite should turn colourless

- Colourless and choking gas evolved

Carbon Dioxide (CO_2):

- Pieces of calcium carbonate (marble chips) with an acid
- Pass the gas through limewater (use of delivery tube)
- Effervescence is observed
- A white precipitate is formed in the limewater

Ammonium (NH_3):

- Ammonium chloride is placed with sodium hydroxide
- Red litmus paper should turn blue (ammonium is basic)
- Colourless and pungent gas evolved

Purification Techniques

Objectives:

1. Obtain pure substance from a mixture (obtaining pure water from a salt solution through desalination)
2. Identify the unknown pure substance, after being separated from a mixture
3. Test for purity of unknown substances (check for purity of food colouring)

Dissolving and Filtering:

- Filter: use of filter paper
- Filtering: separating an insoluble solid from a liquid
- Must be aware of residue and filtrate
- Evaporation to Dryness: separating a soluble solid (solute) from a solution (cannot decompose on heating)
- Evaporation by water bath: similar to evaporation to dryness but no direct heating used (for solutes which are sensitive to heat and easily decompose)

Crystallization:

- Separating a soluble solid from a solution by forming crystals (pure crystals can be used to study the bonding and structure of the substance)
- Many crystals contain water of crystallization
- Water molecules in the crystal give it its shape
- Excessive heating during crystallization would drive off this water and crystals would not form
- A saturated solution is formed (and dried)

Sublimation:

- Separating solids that sublime easily
- Solid substance that becomes gaseous when heated under room conditions without going through liquid state
- Condensed vapour solidifies as a deposit on the surface of the inverted funnel
- For thermally unstable substances
- Ammonium chloride, Dry Ice

Separating Funnel:

- 2 liquids that are immiscible (cannot be easily mixed together) tends to separate in 2 distinct layers
- A tap is at the bottom to remove the denser liquid (that sinks to the bottom)
- Oil and water (which has a higher density)

Distillation:

- Recover a solvent from a solution (with a solid)
- Obtaining pure water from seawater
- Occurs more easily at a reduced pressure (whereby water can be boiled at 70 degrees Celsius)

Boiling chips:

- Anti bumping granules ensures smooth boiling by allowing the formation of smaller air bubbles

Why is the thermometer bulb placed at the end of the flask?

- Measure the temperature of the pure vapour exiting the flask to determine the boiling point of the solvent in the mixture

Why does cool water enter the condenser from the bottom?

- Ensures that bottom of the condenser is the coolest part, and any vapour that did not condense initially would condense here
- Warm water rises upon gaining heat from the vapour, thus water enters through the bottom to allow it to rise to the top upon warming

Fractional Distillation:
- Used to separate liquids in a miscible mixture if they have different boiling points
- Liquids with a lower boiling point are more volatile (vapour will be richer with the substance that has a lower boiling point)
- Glass beads are placed in the fractionating column to provide better separation with a greater surface area for vapour to condense in and from which the liquid can reboil (substance with higher boiling point will condense more easily and trickle back into the mixture)
- Mixture has less of substance with lower boiling point after a period of time
- Used in the separation of different hydrocarbons (kerosene from butane)

Note:
  - Some liquids cannot completely separate (especially if the boiling points are very close to each other)
  - Ethanol and water: distilled ethanol will contain about 4% water
  - Bunsen burners are usually not used: ethanol is highly flammable and no naked flame should head it
  - Another risk: difficult to control the temperature of the flame of the burner to conduct the experiment (requires the stable input of heat for liquid to boil at certain temperature)
  - Water bath or sand is usually used

Chromatography:
- Separating and identifying components present in a mixture (colourless amino acids, plant pigments)
- Able to identify tiny amounts of substance
- Useful in detecting drugs in urine and harmful chemicals of food
- Rf Value: Distance moved by substance/Distance moved by solvent

Reverse Osmosis:
- High pressure used to force impure water through a membrane with tiny pores
- Different from diffusion (not based on random movement of particles based on concentration)
- Only water molecules can pass through, while larger solute particles are left behind
- Purification of water (NEWater)
- Osmotic pressure is applied from the source

Centrifuging:
- Used in medical sciences
- Separate substance of different densities
- Denser substance will stay at the bottom

Solubility Curve:
- Solubility: the amount of solute that will dissolve in 100g of solvent at a certain temperature
- Saturated solution: contains as much dissolved solute as possible at a certain temperature
- Super-saturated: saturated and cooled from a higher temperature (but still saturated at low temp.)
- Curve exists to show the solubility of the solvent to a solute across a range of temperatures

Atomic Structure:

Feature	Proton	Neutron	Electron
Relative Mass	1	1	1/1840
Mass	1.673 x 10^=27	1.675 x 10^-27	9.11 x 10^-31
Electron Volt (EV)	938.27231 MeV	939.56563 MeV	0.51099906 MeV
Charge	+1	0	-1

Atom:
- Basic unit of Matter (smaller: proton, quark)
- Contains protons and electrons (except H+) and neutrons (except H)
- A fixed number of electrons can be accommodated in any shell (calculated by 2n^2)
- Loss of electrons forms cation (with a positive charge)
- Gain of electron forms anions (with a negative charge)
- Cations and anions form mostly to reach a state of structural stability (noble gas configuration)

Noble gases:
- 8 valence electrons to confer stability (causing very low reactivity)
- Atoms and ions have different chemical properties because they have different numbers of electrons
- Noble gases fulfill this naturally (thus do not bond with other atoms, except XF4 and other exceptions)

Electron Cloud Model
- Electrons are not in a fixed shell, but are floating around in empty space (thus there are certain probabilities whereby an electron can appear)
- Different energy levels (of electrons) cannot cross into the region of another "subshell"

Electrons:
- Ability to gain energy to reach an excited state (of a higher orbital level): gains energy from heat and other means (explains the First Law of Thermodynamics)
- Can only jump to a higher level in fixed packets of energy (quanta)
- Falls back to ground state and releases energy in the form of photons/heat and other means

Isotopes: atoms of the same element with the same number of protons but a different number of neutrons
- Same chemical properties as other isotopes (since there are similar number of electrons and protons)
- Different physical properties (mass) due to the presence of different numbers of neutrons
- Radioactive isotopes: give out radiation that is invisible but harmful (due to unstable structure): most are made artificially by humans
- Non-radioactive isotopes (mostly by nature): not harmful and usually naturally occurring
- Relative atomic mass: represents the average mass of one atom, counting the different isotopes and proportions

Mass Spectrometer:
- Sample Inlet
- Ionizer (gaseous sample is bombarded with electrons or laser to create positive ions of the element)
- Mass Analyzer (Ions of the same charge are separated according to their mass to charge ratio by using a magnetic or electric field)
- Detector
- Data Analysis (Mass spectrum in which the mass to charge ratio is plotted against the intensity of the respective ion signals in terms of abundance)

Purpose:
- Able to measure the abundance of a particular isotope of an element
- Able to determine the mass to charge ratio of an ion
- Able to calculate the molar mass (gmol-1) also known as the mass number of an element

Note:
- Large atoms such as plutonium do not have a fixed mass number since most atoms are produced by Man using neutron bombardment (thus the exact abundance cannot be measured)
- Smaller atoms (oxygen) are abundant in nature and their exact mass number can be measured

Ionization Energy:
- Amount of energy that is required to remove one electron from a gaseous atom

- Ability to study the arrangement of electrons (by observing trends

Principles:

- Electrons nearer to the nucleus require greater amounts of energy to remove (because of the attraction between the proton and the electron)
- Electrons from the outermost subshell will be removed first (since the opposite attractions are the smallest)
- Jump in energy required between subshells and shells (with significant difference)
- With each electron removed, the next will require more energy (more protons attracting a fewer number of electrons)
- Measured in log or ln graph (big jump)

Note:

- Nitrogen cannot be ionized 8 times (especially since it has only 7 electrons)

Rules:

Aufbau Principle:

- Electrons must fill up subshells with a lower energy level first
- 1s must be filled before the 2s subshell can be filled (following the actual energy of each shell)

Hund's Rule:

- Every orbital in a subshell is singly occupied with an electron before any one orbital is doubly occupied
- All electrons in singly occupied subshells have the same spin

Pauli's Exclusion Principle (not tested):

- No two electrons can have the same 4 quantum numbers
- Electrons all have half-spins

Octet Rule:

- Most atoms aim to achieve a state whereby there are 8 valence electrons (or 2)

Subshells:

1. Electrons are arranged in "shells": principal quantum numbers (average distance of electron from nucleus
2. Shell consists of s, p, d, f subshells (number of subshells is the shell number)
3. Each shell contains a number of orbitals (and each orbital has two electrons)
4. Electrons fill the subshell with the lowest energy first
5. Subshells are always written in terms of increasing energy levels

Note:

- For energy levels, 4s comes before 3d
- However, when writing the actual electronic configuration for subshells, 3d comes before 4s
- A VERY COMMON MISTAKE ☺

Chemical Bonding and Structure:

Ionic Bonding:
- Giant ionic lattice structure
- Bonding occurs through the **gain** of one electron (for non-metal anion) and **loss** of one electron (for metal-cation)
- Elements from VI, VII and V are usually anions
- Elements from I, II and III (including transition metals) usually become cations
- Chlorine has a stronger pull on the electron of sodium (causing sodium to lose an electron): due to the electronegativity of the two atoms
- Held through strong electrostatic forces of attraction between cations and anions
- Attraction between positive and negative ions
- Octet Rule: atoms try to have 8 electrons in their valence shell for stability

Diagram Drawing:
- For valence electrons, draw an empty shell/no shell for the cation if the question asks for the cation
- Draw a full shell (with appropriate dot and crosses) for anions for valence electrons
- Draw a bracket around the ion and indicate charge

Misconception:
'Ion-pairs' are viewed as molecules
- No molecules in ionic compounds (all ions to attracted to an oppositely charged ion)
- Ions did not 'donate' electrons and are only bonded to one oppositely charged ion
- Ionic bond occurs between ANY two oppositely-charged ion

Box-diagrams are most accurate
- Only accurate for ions with a 1:1 ratio
- Not true for potassium oxide (K2O) which will have a lattice in another shape

Batteries:
- Battery contains electrons in a chamber, which travels towards the cathode (which is usually marked by a '-' sign since cations are attracted to it)
- For Copper (II) sulfate, Cu2+ ions moved to the anode to gain two electrons
- Cu moves towards the anode terminal where it loses electrons to bind with sulfate anions (thus creating a closed circuit)

Physical Properties:
- High melting and boiling points
  - Result of strong electrostatic forces of attraction in the giant lattice structure between anions and cations
  - Usage as refractory heat-resistant materials in industries (line the sides of furnaces at 800 degrees Celsius): industries which require materials with high melting points
  - MgO has a very high temperature of up to 2000 degrees Celsius (an anomaly)
  - Al2O3 is used inside spark plugs
- Ability to conduct electricity in molten and aqueous state
  - Ionic lattice structure can be broken down and ions can move around freely as mobile charge carriers (mobile ions)
  - Solid: ions are held in fixed position and not free to move around (unable to carry a charge)
- Exist as crystals (solids) at room temperature
- Most can be dissolved in water but not in organic solvents
  - Water molecules are polar and can interact with ions (can weaken the electrostatic forces of attraction: causing them to weaken and separate)
  - Hydrophilic and hydrophobic interaction
  - Organic solvents are non-polar and cannot weaken the electrostatic forces of attraction of ions

Covalent Bonds:
- Two atoms share electrons from their valence shell
- Can occur between atoms of the same element (nitrogen and oxygen), or atoms of different elements (like water and carbon dioxide)
- Each atom usually contributes one electron (non-dative) and a covalent bond will constitute two electrons
- Two or more covalent bonds may be formed between two atoms

- Covalent bonds are only formed between atoms that share electrons (other intermolecular bonds not included)
- Octet Rule: covalent bonds usually bring atoms closer to an octet structure (with 2/8/18 or more valence electrons)

Dot and Cross Diagrams:
- Atoms are linked or included in spaces between the two valence electrons (like Venn diagrams)
- Lone pairs are to be shown (Water/Ammonia)

Note:
- Covalent bonds are very strong and difficult to break (covalent substances have high melting points since a lot of energy is required to break bonds between atoms in the covalent substance: for diamond)
- Covalent substances have low melting points due to the weak intermolecular forces of attraction

Physical Properties:
- Unable to conduct electricity
    - ➢ Does not have any delocalized ions or mobile charge carriers at any state (since all electrons are shared and occupied)
    - ➢ Made up of neutral molecules
    - ➢ May be able to conduct electricity only at under certain circumstances (when the covalent bonds in water molecules are split or when molecules are dissolved in water to become acids)
- Unable to dissolve in water but can dissolve in organic solvents (for simple molecular structures)

Simple Molecular Structures:
- Usually not large enough in a particular structure
- Contains the intermolecular forces of attraction
- Examples include: Carbon Dioxide or Oxygen

Physical Properties (not including those relevant for all covalent substances):
- Gas and liquid states at room temperature
- Very low melting and boiling points
- Exception: Iodine can exist as solids at room temperature
- Unable to dissolve in water but in organic non-polar solvent (eg. Iodine is only partially soluble in water but dissolves well in hexane)

Macromolecular Structures:
- Occurs when a large number of atoms are joined together by covalent bonds
- Extends throughout network structure, mostly with no intermolecular forces of attraction (except graphite)
- Example include: fullerene, diamond, sand, graphite (some of which may not be made up of atoms of the same element)
- Includes also: polymers and internal proteins (with nucleic acids)

Physical Properties (not including those relevant for all covalent substances):
- Usually not able to conduct electricity since it is made up of neutral atoms with no mobile charge carriers (except for graphite)
- Not soluble in any solvent (Sand and diamond do not dissolve in polar or non-polar solvent)

Diamond:
- Melting point of 3700 degrees Celsius
- 3.5g/cm^3 density
- Appearance: colourless, transparent crystals
- Hardness: hardest natural substance: structure consists of carbon atoms in strong covalent atoms arranged in tetrahedral units
- Electrical Conductivity: does not conduct electricity since all valence electrons are involved in bonding with other carbon atoms (no delocalized electrons or charge carriers to move throughout the structure and carry current to conduct electricity)

Graphite:
- Melting point of 3300 degrees Celsius
- 2.2g/cm^3 (with air spaces in the spaces between each layer)
- Black shiny powder

- Hardness: soft and can be used as lubricant (with parallel layers of carbon atoms, the weak intermolecular forces between layers allow them to slide against each other)
- However, graphite is usually hard in one of the dimensions
- Electrical Conductivity: able to conduct electricity since each carbon atom is only bonded to three others, and each atom has one valence electron not bonded (delocalized electron can move freely along carbon atom layers to carry current and conduct electricity)
- Distance between each layer is 2.5x the distance between atoms

Sand:
- Made of silicon dioxide (which can be combined into a giant structure)
- Strong covalent bonds that extend throughout the network structure
- High melting and boiling point
- Very hard and non-conductor of electricity in any state

Polymers:
- Considered covalent bonds (and a macromolecule)
- Consists of large molecules and not in a network or tetrahedral structure
- Do not have a high boiling point
- Includes plastics, hydrocarbons

Dative Bonding (Co-ordinate Bond):
- Covalent bond in which the shared pair of electrons is provided by only one of the bonded atom
- One atom is the donor, the other is the acceptor
- Same characteristics as a covalent bond once dative bond is formed
- Examples: Hydronium ion (H_3O^+) or Ammonia+Aluminum Chloride in NH_3AlCl_3 (where N donates to Cl)
- Complex example: $CuCl_4^{2-}$ (where all Cl atoms donate to the Cu)

Polar Covalent Bonds:

Electronegativity:
- Measure of tendency of an atom to attract a bonding pair of electrons
- O and H form bond (O pulls the shared electron towards itself, creating partial negative charge-polar covalent bond)
- Pauling scale: most commonly used as measure of electronegativity of an atom (Fluorine with 4.0 and Francium with 0.7)
- Atoms at the top of the scale in the Group attract electrons more strongly, while atoms nearer to Group VII have a stronger pull as well

Difference in electronegativity
- Same electronegativity=likely to be id-id bond (electrons on average are found half way between the two atoms: usually a pure covalent bond)
- Difference in electronegativity=likely to be pd-pd or hydrogen bonding (one atom will attract the electron pair more, creating a partial charge)
- Ionic: one atom loses all control of the electron pair (thus the other has complete control over electrons)
- 0-0.5: Non-polar covalent (0-0.5 refers to the difference in electronegativity of the two atoms)
- Non-polar covalent solvents dissolve in organic solvents easily but not in water
- 0.5-2: Polar covalent
- Polar covalent solvent can dissolve in organic solvents, but some can dissolve in water
- 2-4: Ionic
- Usually dissolves in water and not organic solvents

Polarizability:
- Negative ion that readily undergoes a large distortion is said to be highly polarizable
- Positive ion that can cause large distortion=high polarizing power
- Trend 1: larger negative ion is more highly polarizable (due to the distance of the electron from the nucleus and the attraction between the proton and the electron)
- Trend 2: positive ion has a strong polarizing power if it is small and highly charged (eg. Aluminum) due to a stronger pull on the electron

Dipoles:

- Separation of opposite charges
- Size of dipole is known as dipole moment (measured in debye or coloumb metre)
- Dipole moment can cancel each other out (Carbon Dioxide in linear form-where there is no polarity in the molecule)
- Bond angles and lone pairs decide if dipoles are formed
- Water is a polar molecule (wince H2O lie at an angle with the presence of a lone pair)
- Methane is a non-polar molecule which ammonia is a polar molecule (based on lone pair)

Instantaneous-Dipole Induced Dipole (Van der Waals forces):

- Electrons are in continual motion (since electron charge cloud moves along with the movement of the electron)
- Molecule is not symmetrical due to electron movement (one side will be more negative than another)
- Instantaneous dipole (at that moment) can induce a dipole on another molecule (but are short lived since electrons move again and another dipole moment is formed(
- Factor 1: number of electrons in the molecules (greater number=greater id-id interaction)
- Factor 2: Surface area of molecule (greater surface area= greater id-id interaction)

Permanent Dipoles:

- For polar molecules
- Align such that delta positive molecule is near the delta negative portion of another molecule
- Electrostatic forces of attract between the two ends (leads to pd-pd bond)
- Stronger than id-id forces (have a higher temperature for melting and boiling than id-id)

Hydrogen Bonding:

- One molecule must have a hydrogen bond bonded to N, O or F
- Electrostatic force of attraction between H partial positive and N/O/F partial negative
- Hydrogen atom has no inner core electrons (electrons are attracted by atom hydrogen is covalently bonded to) and the partially exposed hydrogen proton interacts strongly with the lone pair of an electronegative atom leading to a hydrogen bond

Why is the density of ice lower than water?

- Ice is a highly ordered 3 dimensional shaped solid (due to hydrogen bonding)
- Order creates a very open structure and prevents molecules from getting too close to another
- Density decreases due to the free space (ice has a density of 0.92g cm-3 while water is 1.00gcm-3)

Metallic Bonding:

- Outermost electron in each atom are shared equally by all the metal atoms
- Delocalized electrons can conduct heat and electricity
- Sea of freely moving delocalized electrons: metal atoms are held by electrical interaction between positively charged ions and negatively charged delocalized electron
- Electrons are not evenly distributed (but tend not to congregate in an area due to electron repulsion)
- Malleable and ductile (atoms arranged in layers that can slide over each other easily)

Physical Properties:

- High melting and boiling temperatures (electrostatic forces of attraction between ion and electron): more heat energy required to overcome these forces
- All conduct electricity in solid and liquid states (metallic lattice has positive ions surrounded by sea of electrons which act as charge carriers)
- Not soluble in any solvent (metals may react with water though-a chemical change and not a physical change)

Stoichiometry and the Mole:

Formulas:
- Molecular formula shows all the atoms in one molecule
- Empirical formula: simplest ratio shown
- Possible for two different substances to have the same empirical formula (glyceraldehyde and glucose)
- Structural formula: shows how all the atoms are joined (in a diagram)

Constructing equations:
- Take note if a word/chemical equation is required
- Take note if state symbols are required
- State symbol: used to describe if reactant or product is in the solid, liquid, aqueous or gaseous form
- Insoluble in water, usually a solid
- Soluble in water: aq state symbol

Solubility:
- All nitrate salts
- Most chloride salts (except lead II and silver chloride)
- Most sulfate salts (except lead II, barium and calcium sulfate)
- Carbonates of sodium, potassium and ammonium and most other carbonates
- Hydroxides and oxides of sodium, potassium and ammonium and most other hydroxides and oxides
- All acids and alkalis

Atomic Masses:
- Relative Atomic Mass: the ratio of the average mass per atom of the naturally occurring form of an element to one-twelfth the mass of an atom of carbon-12
- Mr (relative molecular mass) has no unit (relative quantity)
- Molar mass: has the unit of gmol-1 (mass of 1 mol of the substance)

Avogadro's Constant:
- Measuring the number of atoms in 1 mole of a particular substance
- Measured to be about 6.02 x 10^23

Formulas for Mole Concept:
- No. of Moles x Molar Mass = Mass of Sample
- No. of Moles x Avogradro's Constant= No. of Particles
- No. of Moles x 24.0dm3 = Gas Volume

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.

Limiting Reactant:
- Reactant is completely used up=Limiting
- Reactant that is not completely used up=Excess Reactant
- Left over=mass of reactant that is unused in the reaction
- Usual question: identify limiting reactant before finding left over and mass of product

Percentage Purity
- [Mass of pure substance/Mass of given sample] x 100%
- Percentage Yield = [Actual Yield/Expected Yield] x 100%
- Percentage Yield calculated: since chemical is usually lost in the reaction

Mole Ratio:
- Important in calculating the number of moles of any substance (1:2/1:40)

Molar Volume of Gases:
- Equal volumes of all gases under the same conditions of temperature and pressure will contain the same number of molecules
- One mol of any gas under the same condition will have the same volume
- Value is interchangeable with no. of moles (mole ratio is the volume ratio)
- 1000cm3=1dm3 (all values must be converted to dm3 before any further calculations)
- Only for use between GASES

Misconceptions:
- Amount vs Mass (Amount=No. of Mol, while Mass=g)
- Leave exact value (usually round off to 4sf in intermediate step but 3sf in final solution)

Exam tips and techniques:

Common Misconceptions:
Use of “heavier”:
- 1 gram of Bromine Gas will be lighter than 1000g of Hydrogen Gas
- Use “one molecule” of Bromine Gas is heavier than “one molecule” of Hydrogen Gas (due to Mr)

Octet Rule:
- Usually states that atoms combine in a way such that atoms attain 8 electrons in outer shell
- Notable exceptions: Boron trifluoride (with only 6 valence electrons) or sulfur hexafluoride
- Also, octet rule not necessarily means 8 electrons (but also 2 and 18, depending on period)

Ionic compounds:
- Usually when metal reacts with non-metal, an ionic compound is formed
- Lead chloride has low m.p. of -15 degrees Celsius (simple covalent structure) and does not conduct electricity

Distinguishing Elements in Diagrams
- Provide a legend (use of shading)
- Use different forms of shading
- Do not use shape and size as a basis (if possible-not distinct enough)
- Do not use pens of different colours as much as possible (to shade)

Periodic Table of Elements:
- Identify the distinction between metals and non-metals (draw the staircase)
- Mark out tricky elements (Ag-which can only be in the form of Ag+, and Zinc, which can only be 2+)
- Write down –ide, -ite, and –ate (for sulfate or nitrate)
- List down common acids (nitric/sulfuric/carbonic)