- **1.1 Introduction to the particular nature of matter and chemical change**
  - **States of matter**
    - 3 states: solids, liquid, and gas.
    - Solid definite volume & shape which can be affected by Temp.
      - Thermal expansion & thermal contraction.
    - Liquid definite volume & variable shape, slightly affected by Temp
    - Gas variable volume & variable shape
      - Gases will spread out evenly in containers, known as diffusion.
      - Greatly affected by Temp
  - **Elements**
    - Elements are the simplest substances & are each composed of a single type of atom.
    - Elements can't be split or decomposed into simpler substances by a chemical reaction.
    - Separations like metals and non-metals, based on chemical & physical properties.
    - Metals & noble gases exist as atoms while many non-metals exist as molecules (atoms bonded with one another).
      - P_4 & S_8 are molecules like O_2, N_2, and Cl_2.
    - O_2, N_2, and Cl_2 are diatomic.
    - Allotropy is the existence of two or more crystalline forms of an element.
      - Called Allotropes.
      - Allotropes exist when there is more than one possible arrangement of bonded atoms.
        - E.g Carbon exists as Diamond, graphene, graphite, and C-60.
        - E.g Oxygen exists as O_2 & O_3.
  - **Compounds**
    - The formation of a compound (two elements chemically bonded together) is known as synthesis.
      - Usually this is an exothermic reaction (like other reactions where bonds are formed).
      - E.g iron & sulfur is heated, energy released when iron (II) sulfide, FeS, is formed.
    - Unlike compounds, mixtures aren't chemically bonded and can therefore be separated by physical methods.
      - Compounds have different chemical properties to their constituent elements. Mixtures retain them.
  - **Molecular kinetic Theory**
    - In crystalline solids, particles are in a lattice
    - In solids the particles will vibrate in a fixed position, and as Temp increases they will vibrate more & more.
      - This is due to the increase in Temp.
      - The less E_k they have, the stronger the force of attraction between them.
  - **Changes of state**
    - **Heating and cooling curves**
      - When a substance is changing state, e.g ice to water, not all of the ice will have turned into water instantly.
        - There will be approximately a 1:1 ratio of $H_2O_{(l)}$ to $H_2O_{(s)}$ molecules.
        - The same occurs when water is at 100°C.
      - In a heating & cooling curve changes in state are represented with a straight line as all the energy is dedicated to breaking the intermolecular bonds.
      - **Atom economy in chemical reactions**
        - Atom economy examines the theoretical potential of a reaction, by dividing the quantity of desired product by initial amount of reactants.
          - $\%\text{ atom economy} = \dfrac{\text{atomic mass of all utilized atoms}}{\text{atomic mass of all reactants}}$
  - **Mixtures**
    - Two types of atoms present which aren't chemically bonded to one another.
    - Compounds will always contain the same proportion (by mass) of each element. While mixtures can have any proportion of each element.
    - Alloys are mixtures of metals and other elements that have been melted together and then allowed to solidify.
  - **Types of mixtures**
    - Mixtures of gases are said to be homogeneous since they have a uniform or constant composition.
      - Homogeneous solutions are solutions which have the same concentration at any point in the solution.
    - Gases diffuse quickly in containers because they move quickly due to high kinetic energy.. also due to the large spaces between molecules.

- this process is much slower with liquids due to lower kinetic energy.
- Particles of water (the solvent) will collide with particles of the substance being dissolved (the solute).
    - When they collide with one another, they'll attract.
- Water molecules pull off and interact with the solute particles from the solid solute.
    - The water molecules surround the solute particles.
    - As the water molecules move throughout the solution, the solute particles will spread out.
- In a heterogeneous mixture the composition isn't uniform throughout the mixture.
    - E.g water & oil will make two different layers. They're said to be immiscible.
- Water & ethanol will form a uniform layer.
    - They're a homogeneous mixture.
- On a macroscopic level, matter can be classified into mixtures or pure substances.
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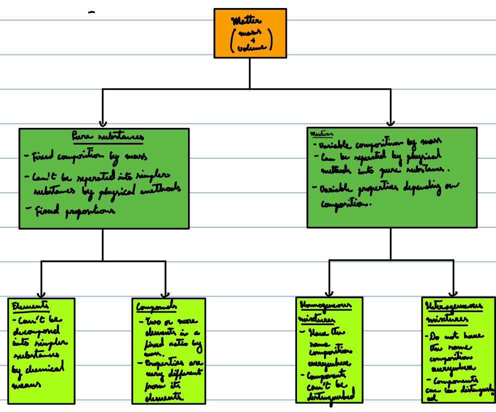


- Chemical formulas
    - A polyatomic or compound ion is an ion that contains more than two covalently bonded atoms with an associated charge.
    - Since ionic compounds are neutral, the ions which are covalently bonded must have opposite charges which cancel one another out.
    - The value of the charge of an element can be determined by its place on the periodic table.
        - If the element is closest to the noble gas group by moving to the right, then it will be negative showing that they need to gain electrons to be a noble gas.
            - An example of this would be the ion of oxygen, oxide which has a charge of negative one (O^-).
        - If the element is closest to the noble gas group by moving to the left, then it'll be a positive ion because of the fact that it'll lose electrons to become a noble gas.
            - E.g Copper who's ion is: Copper (II) Cu^{2+}.
- Ionic equations
    - When a soluble ionic substance is dissolved in water, the ions seperate.
        - E.g $BaCl_{2(s)} + (aq) \longrightarrow BaCl_{2(aq)} \longrightarrow Ba^{2+}_{(aq)} + 2Cl^-_{(aq)}$
    - The Barium & Chloride ions undergo their characteristic reactions regardless of which other ions are present in the solution.
    - Solubility of common salts:
        - All sodium, potassium, and ammonium salts are soluble.
        - All nitrates are soluble
        - All chlorines are soluble except silver chloride & lead (II) chloride.
        - All sulfates are soluble except calcium sulfate, barium sulfate and lead (II) sulfate.
        - Sodium, potassium, and ammonium carbonates are soluble but all other carbonates are insoluble.
    - Ions are removed from solution by the following process:
        - Formation of an insoluble percipitate.
        - Formation of molecules containing only covalent bonds.
        - Formation of a new ionic chemical species.
        - Formation of a gas.
- 1.2 The mole concept
    - The mole concept and the Avogadro constant
        - 12 grams of C-12 atoms and 1 gram of H-1 atoms both contain the same number of atoms.

- One mole of a substance contains $6.02 \cdot 10^{23}$ particles of the substance.
- Relationship between the amount of a substance and the number of particles:
    - $\text{amount of substance (mol)} = \dfrac{\text{number of particles}}{6.02 \cdot 10^{23}}$
- $\text{Avogadro's constant} = \dfrac{\text{volume of one mole of atoms}}{\text{volume of one atom}}$
    - $\text{Volume} = \dfrac{nM_m}{\rho} = \dfrac{\text{relative molecular mass}}{\text{density}}$
- Questions:
    - Calc number of molecules of H_2O in 0.01 mol of water:
        - $0.01 = \dfrac{\text{number of particles}}{6.02 \cdot 10^{23}}$

    number of particles $= 6.02 \cdot 10^{21}$ ✓

    - $\text{mol} = \dfrac{9 \cdot 10^{23}}{6.02 \cdot 10^{23}}$
      $= 1.495$
      ≈ 1.5 mol ✓
    - HNO_3, number of atoms of oxygen are: $6.02 \cdot 10^{23} \cdot 3 = 2.7 \cdot 10^{24}$ atoms

- Formulas
    - Relative atomic mass, relative formula mass, and molar mass
        - The relative atomic mass of an element is how many times greater the average mass of atoms of that element is than one twelfth the mass of C-12 atoms.
            - $\text{relative atomic mass } (A_r) = \dfrac{12(\text{average mass of one atom of the element})}{\text{mass of one atom of carbon-12}}$
                - A_r of Hydrogen $= \dfrac{12(1.01)}{12} = 1.01$
        - The relative molecular mass is the sum of the relative atomic mass of all the atoms in one molecule.
            - E.g M_r of oxygen is 32 and the A_r is 16, the M_r is 32 because oxygen exists as diatomic molecules (O_2).
            - It has no units.
- Molar mass
    - The molar mass (M) is the mass of one mole of any substance where C-12 is assigned a value of exactly $12 gmol^{-1}$.
    - $\text{Amount of substance (mol)} = \dfrac{\text{mass (g)}}{\text{molar mass } (gmol^{-1})}$
        - The molar mass of an atom or ion is equal to its relative atomic mass expressed per mole.
    - E.g molar mass of $Pb(NO_3)_2$ is $207.2 + 2(14.01 + 3(16)) = 331.21 gmol^{-1}$
- Calculating quantities
    - $\text{mol} = \dfrac{g}{gmol^{-1}} = \dfrac{54}{33.01} \approx 1.64$ mol — $\text{mol} = \dfrac{500g}{40} = 12.5$ mol — $\text{mol} = \dfrac{0.18}{33.01} \approx 0.00545 = 5.45 \cdot 10^{-3}$ mol

- Molecular and empirical formula
    - The empirical formula is the simplest possible integer from the compound.
        - Eg $C_6H_{12}O_6$ is CH_2O (empirical formula).
        - Carbon = 39.13%, Oxygen = 52.17%, and hydrogen = 8.7%. Carbon = 3.27, Oxygen = 3.26, Hydrogen = 8.61 $= \dfrac{8.61}{3.26} = 2.64$ = $C_3H_8O_3$ — 1:1:2.6 ratio = 1:1:3 ratio — Why multiply by 3?
        - $Pb = \dfrac{41.4g}{207 gmol^{-1}} = 0.2$, Oxygen $= \dfrac{3.2g}{16 gmol^{-1}} = 0.2$, P = 0.2 mol, O = 0.2 mol = 1:1 ratio $\boxed{PbO}$ ✓
    - The molecular formula represents the actual number of atoms in a molecule of a simple covalent substance.
    - Another situation where empirical formula must be derived is the empirical formula of a hydrated salt where the water of crystallization can be removed without the anhydrous salt undergoing decomposition.
        - 12.3 grams of hydrated magnesium sulfate, $MgSO_4 \cdot xH_2O$, gives 6 grams of anhydrous magnesium sulfate, $MgSO_4$.
            - $12.3 - 6 = 6.3g$

            $MgSO_4$: $\dfrac{6g}{120 gmol^{-1}} = 0.05$ mol — H_2O: $\dfrac{6.3g}{18 gmol^{-1}} = 0.35$ — $MgSO_4 7H_2O$

            1:7
- 1.3 Reacting masses and volumes
    - Calculating theoretical yields

- Steps to solve stoichiometric problems
  - If necessary, provide a balanced equation
  - Convert mass/volume in moles
  - Calculate amount of required product (using coefficients)
  - Convert the amount of product to appropriate units of mass.
- $CaCO_3 \longrightarrow CaO_{(s)} + CO_{2(g)}$
  - $\text{moles} = \frac{g}{g\,mol^{-1}}$
    $= \frac{2.5g}{100g\,mol^{-1}}$
    $= 0.025\ mol$
  - 1 mole of CaO & 1 mole of CO_2 meaning 0.025 mol of $CaO_{(s)}$ & 0.025 mol of $CO_{2(g)}$.
  - $g = \text{moles} \cdot g\,mol^{-1}$
    $= 56 \cdot 0.025$
    $g = 1.4g$ of $CaO_{(s)}$
- $N_2 + 3H_2 \longrightarrow 2NH_3$
  $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$
  $2NO + O_2 \longrightarrow 2NO_2$
  $2H_2O + 4NO_2 + O_2 \longrightarrow 4HNO_3$
  - Overall 1 mole of N_2 produces 2 mol of HNO_3.
  - $N_2 = \frac{56g}{28g\,mol^{-1}}$
    $= 2\ mol$
    $= 4$ mol of N_2
  - $g = 4 \cdot 63$
    $= 252g$ of nitric acid

- **The limiting reactant and the reactant in excess**
  - For the following reaction: $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$.
    - The reaction only requires one mole of oxygen.
      - meaning that only one mole of oxygen will be left over when the reaction is complete.
    - The amount of water obtained is determined by the amount of reactants that is completely consumed during the reaction.
      - The one left over is in excess.
  - Worked example
    - 4.8 grams of Mg & 4.8 grams of sulfur
      - $\frac{4.8}{24} = 0.2\ mol$ $\frac{4.8}{32} = 0.15\ mol$ $\boxed{Mg + S \longrightarrow MgS_{(s)}}$
        Limiting reactant is sulfur
      - $g = 0.15 \cdot (56)$
        $= 8.4g$ of MgS formed
        $g(\text{Mg unreacted}) = 0.05 \cdot 24$
        $= 1.2g$

- **Percentage and experimental yields**
  - The theoretical yield is when all the limiting reactant reacts, while experimental yield is the actual product actually obtained.
    - The experimental yield is always less than the experimental one, due to:
      - side or competing reactions.
        - Competing reactions are reactions which take place that aren't the desired reaction. E.g $C_{(s)} + O_2 \longrightarrow CO_{2(g)}$ will have a side reaction of $2C_{(s)} + O_{2(g)} \longrightarrow 2CO_{(g)}$.
      - The reaction is reversible and reaches equilibrium.
      - Impurities present in the reactants
      - mechanical losses/physical losses of the reactants or products
        - Stuck in glassware or filter paper
  - Percentage yield can be calculated by:
    - $\text{Percentage yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \cdot 100$

- **Percentage Purity**
  - It is the percentage of a specified compound or element in an impure sample.
  - Calculated with following formula:
    - percentage purity = mass of pure substance in a sample ·100

mass of sample

- Reacting volumes of gas
  - A fixed amount of gas will always react with an exact volume to form a product
    - Given by formula: $V \propto n$
- Molar volume of gas
  - Gas molar volume has the value of 22.7 dm^3 at 0°C and 1 atmosphere.
    - Volume of gas = mol · 22.7 dm^3
    - $\text{mol} = \frac{\text{volume of gas}}{22.7\,dm^3}$

- Relationship between temperature, pressure, and volume of a gas
  - As volume decreases pressure will increase
  - As temp increases, pressure decreases as volume increases
  - As volume decreases temp increases
  - Gas density
    - The mass of the gas divided by its volume.
    - Increase in pressure will lead to decrease in volume and increase in density.
    - If temperature increases, then volume decreases and density decreases.
- Kinetic theory of gases
  - Kinetic theory of gases makes the following assumptions
    - The individual molecules or atoms of a gas have a negligible volume compared to the volume of the container.
    - No attractive or repulsive forces between atoms or molecules except for when they collide with one another or with the container.
    - Collisions are perfectly elastic and are the reason for pressure in the container.
    - The average kinetic energy of the molecules or atoms in the gas is directly proportional to the absolute temperature.
      - E_k of the gas molecules is given by $k_E = \frac{1}{2}mv^2$.
  - This is a description of ideal gases. Real gases act as ideal ones when there are high temperatures & low pressures. Although no gas really acts as an ideal one as the first two assumptions aren't quite true.
- The gas laws
  - Boyle's law
    - $P \propto \frac{1}{V}$ or $P \propto V^{-1}$ at constant temperature
    - Can be used to calculate the new pressure or new volume, given by the following formula: $P_1 \cdot V_1 = P_2 \cdot V_2 \rightarrow \frac{P_1}{P_2} = \frac{V_2}{V_1}$.
    - Worked examples
      - 350 cm^3 = V_1, P_1 = 103 kPa, P_2 = 150 kPa

$$\frac{P_1}{P_2} = \frac{V_2}{V_1}$$
$$V_2 = \frac{P_1 V_1}{P_2}$$
$$= \frac{(103000)(350)}{(150000)}$$
$$V_2 = 240.3\,cm^3$$

  - Charles' law
    - $V \propto T$, at constant pressure
    - Used to calc new temp and volume: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$.
      - Temp in Kelvin
    - Worked example
      - V_1 = 4.5 dm^3, temp_1 = 300K, temp_2 = 350K

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
$$V_2 = \frac{(4.5)(350)}{300}$$
$$V_2 = 5.25\,dm^3$$

  - The pressure law
    - At constant volume, $P \propto T$
    - $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

    - Question: 10 dm^3 = V_1, P_1 = 97 kPa, T_1 = 25°C, P_2 = 101325 Pa

P_1 . P_2

$$\frac{}{T_1} = \frac{}{T_2}$$

$$T_2 = \frac{(101325)(25+273.15)}{97000}$$

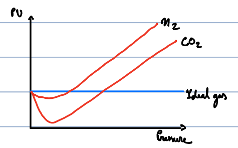
$$= 311.44\,K$$

- Equations of state (combining gas laws)
    - $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \longrightarrow \frac{PV}{T}$ known as equation of state

- Ideal gas equations
    - $PV = nRT$
        - P = pressure, V = volume (m^3), n = number of moles, R = gas constant ($8.315\,JK^{-1}mol^{-1}$), and T = temperature (Kelvin)
    - $1\,dm^3 = 10^{-3}\,m^3$ & $1\,cm^3 = 10^{-6}\,m^3$.
    - $PV = nRT = \frac{mRT}{M_r} \longrightarrow M_r = \frac{mRT}{PV}$
    - $d = \frac{m}{V} \rightarrow V = \frac{m}{d}$
    - $M_r = \frac{dRT}{P}$
    - Question: $M_r = \frac{dRT}{P}$

$$= \frac{(2.615)(8.31)(298)}{(101.325)}$$

$$= 63.9$$

    - 

- Solutions
    - Water is a solvent and the substances dissolved in it are solutes.
    - The mixture is called a solution.
    - concentration = [ ].
    - One mole of solute dissolved in water and volume of solution made up $1\,dm^3$, resulting solution = molar solution.
    - Concentration of a solution is the amount of solute (moles) contained within one cubic decimeter (dm^3).
        - concentration = $\frac{\text{mole}}{\text{volume}} \rightarrow c = \frac{mol}{V}$
    - Parts per million (PPM) is used for very dilute substances.
        - ppm = $\frac{\text{mass of concentration in solution}}{\text{total mass solution of}} \cdot 10^6$

    - Dilutions of acids
        - Dilution is done by adding a strong acid to water.
            - Water should never be added to a strong acid as it would form an extremely concentrated solution of acid initially. Meaning a lot of heat would be released that the solution may boil violently, splashing concentrated acid because the reaction is exothermic.
        - The solute in the water remains unchanged when a solution is diluted: $M_1 V_1 = M_2 V_2$
            - M = concentration, V = volume
            - Question: $V_1 = 25\,cm^3$, $C_1 = 5\,mol\,dm^{-3}$, $1.5\,mol\,dm^{-3} = C_2$

$$M_1 V_1 = M_2 V_2$$

$$V_2 = \frac{(5)(0.025)}{1.5} \cdot 1000$$

$$V = 83.3\,cm^3 \checkmark$$

    - Volumetric chemistry
        - A solution of known concentration is called a standard solution.
        - In volumetric chemistry titration is when a small amount of known concentration (titrant) is added to a known volume of an unknown concentration until a colored gel shows that neutralization has occurred.
    - Acid-base titrations (titration is when a solution of known concentration is used to determine the concentration of the unknown solution)
        - They're reached in the presence of a suitable acid-base indicator, including:
            - Determining the concentrations of solutions
            - Determining the percentage purity or molar mass of an acid or base
            - Deducing the equations for a neutralization reaction

- Determining the amount of water of crystallisation in a hydrated salt.

- Questions: $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$, 0.05 M NaOH, 25 cm^3 of 0.2 M HCl

$\frac{250}{1000} dm^3 \cdot 0.2\,M = 5 \cdot 10^{-3}$ mol HCl   1:1 ratio between NaOH & HCl therefore amount of NaOH = 0.05 mol dm^{-3}

volume of NaOH $= \frac{5 \cdot 10^{-3}\,mol}{0.05\,mol\,dm^{-3}}$ $\boxed{V = \frac{mol}{c}}$

$= 0.1\,dm^{-3}$

$= \underline{100\,cm^3}$

Primary standard solutions

- titrations often involve a primary standard solution. ?
  - the concentration may have been determined by titrations with another primary standard solution.
- the concentration of primary standard solutions don't change with time.
- If a substance must be weighed accurately enough for use in preparing a primary standard solution, the following criteria must be met.
  - the substance must be in a state of high purity
  - the substance must not be volatile, or some of it will be lost
  - Can't react with H_2O, CO_2, or O_2.

Back titrations

- For back titration, a known excess of one reagent A is allowed to react with an unknown amount of a reagent B.
  - At the end of the reaction, the amount of A that remains unreacted is found by titration.
- A typical acid-base titration, a quantity of a base is added to an excess amount of acid. All of the base will react & only part of the acid will react.
  - the remaining acid is then titrated.
  - therefore showing the amount of base can be calculated.
- Questions: 4.08g MgO v = 100 cm^3 HCl, c = 2M HCl) 19.7 cm^3 of 0.2M NaOH

a) $MgO_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)}$

$NaOH_{(s)} + HCl_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

b) 1:2 ratio   MgO mol = 0.1 mol   HCl mol = cv

$= (2\,mol\,dm^{-3})(\frac{100}{1000})$

$= 0.2$ mol

c) $\frac{19.7}{1000} \cdot 0.2 =$ mol

mol = 0.00394 mol

d) $0.2 - 0.00394$

$= 0.196$ mol

e) $\frac{0.196}{2} = 0.098$ mg

$0.098\,mol \cdot 40.3\,g\,mol^{-1} =$ mass

mass = 3.95g

$\frac{3.95}{4.08} = \underline{96.8\%}$

Redox titrations

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