3.2 Chemistry: Stoichiometry, Acids and Bases, Ionic Reactions

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
- Structural formula: shows how all the atoms are joined (in a diagram)

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

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)

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)

Acids and Bases:

- Acid: an acid is a substance that ionizes in water to produce hydrogen ions
- Alkali: an alkali is a soluble base that ionizes in water to produce hydroxide ions

Common Acids:

Common Name	Chemical Name	Formula	Description
Hydrochloric Acid	Aqueous Hydrogen Chloride	HCl	Strong acid, monobasic
Nitric Acid	Aqueous Hydrogen Nitrate	HNO3	Strong acid, monobasic
Sulfuric Acid	Aqueous Hydrogen Sulfate	H2SO4	Strong acid, dibasic
Carbonic Acid	Aqueous Hydrogen Carbonate	H2CO3	Weak acid, dibasic
Acetic acid (Vinegar)	Aqueous Ethanoic Acid	CH3COOH	Weak acid, monobasic

Common Alkalis:

Common Name	Chemical Name	Formula	Description
Caustic Soda	Sodium hydroxide	NaOH	Strong alkali
Caustic Potash	Potassium hydroxide	KOH	Strong alkali
Slaked lime	Calcium Hydroxide	Ca(OH)2	Strong alkali (but only slightly soluble in water)
Ammonia solution	Aqueous Ammonia	NH3	Weak alkali

Chemistry of Acids:

1. Acid reacts with Base to form Salt and Water (Neutralization)
2. Acid reacts with Carbonate to form Salt, Water and Carbon Dioxide
3. Acid reacts with (some) Metal to form Salt and Hydrogen
4. Acid reacts with Sulfite to form Salt, Water and Sulfur Dioxide

Chemistry of Bases:

1. Base reacts with Acid to form Salt and Water (Neutralization)
2. Base reacts with Ammonium salt to form Salt, Water and Ammonia
3. Alkali reacts with (some) Salt solutions to form Salt and insoluble Hydroxides (precipitation reactions)

Strength and Basicity of Acids:

Strength:

- Extent of dissociation/ionization
- A strong acid has acid molecules that are completely ionized in an aqueous solution
- A weak acid has molecules that are partially ionized in aqueous solution

Concentration:

- Measurement of the amount of solute in a solvent
- The industrial benchmark for a concentrated acid is 2 moldm-3

Basicity:

- Dependent on the number of H atoms in a molecule that are able to ionize to form H+ ions
- Hydrochloric acid is able to provide 1 hydrogen ion per molecule (monobasic)
- Sulfuric acid is dibasic while phosphoric acid is tribasic
- Not all the hydrogen atoms in a molecule ionizes to form H+ ions (in ethanoic acid, CH3COOH, only the H in the carboxyl functional group is dissociated)
- The basicity of an acid does not represent the strength of an acid.

Conditions for dissolving an acid:

- Requires water (or a polar solvent capable of dissociating the acid)
- The acid must dissolve in the water and dissociation must occur to form H+ ions (before it displays its acidic properties)
- HA + H2O → H3O+ + A- (where the acid molecule has been ionized, with A representing a molecule bonded to the hydrogen ion)

Writing Formulae of Salts (Neutralization):

- When an acid reacts with a base, the formula of the salt formed is derived by replacing the H+ of the acid and the metal ion of the base
- This replacement of the H+ must obey the rule of balancing charges (when HCl reacts with MgO, only MgCl2 can be formed, since one Mg2+ replaces 2 H+)

- For dibasic and tribasic acids, since there are 2H+ that can be replaced, 2 possible salts can form
    - When H2SO4 reacts with NaOH, if one H+ is replaced, the salt formed is NaHSO4. If two H+ are replaced, then Na2SO4 is formed.

<u>pH of a Solution:</u>

- pH of a solution measures the extent of alkalinity or acidity of a solution
- pH of a solution is given as the negative logarithm to base ten of the molar hydrogen ion concentration (pH=-lg(H+))=
- pH values can be above 14 and lower than 0: only in cases of very high concentration
- Important: an aqueous solution always contains H+ and OH- ions, due to the small extent of partial dissociation of water molecules (H2O → H+ and OH-)
    - If a solution is acidic, [H+] > [OH-] and the pH < 7
    - If a solution is neutral, [H+] = [OH-] and the pH=7
    - If a solution is alkali, [H+] < [OH-] and the pH > 7
- pH of a solution can be measured by:
    - Chemical indicators which show different colours at different pH values (Universal Indicator and litmus paper)
    - pH meter/sensor: an electrical device more accurate than indicators and must be dipped into the solution

<u>Acid-Base Titration Curve:</u>

- Acid-Base Titration: Used to determine the concentration of an acid or base. This is done by exactly neutralizing the acid/base with a base or an acid with known concentration
- When acid and base are mixed in exactly the right proportions for neutralization, the equivalence point is reached.
- When the indicator changes colour, this is the "end point" of a titration
- NOTE: CHECK whether the acid or base is the original solution present (start at either low or high pH)

Study Strong Acid-Base, Weak Acid-Strong Base and Strong Acid- Weak Base Graphs
Weak Acid-Weak Base:

- Not often titrated since the colour change shown with the indicator is often quick, and therefore difficult for the observer to see the change of colour
- The graph does not have a sharp plunge in the pH and the equivalence point is difficult to determine

<u>Types of Oxides:</u>

Metal Oxides:

- Basic Oxides: react with acids to form salts
    - Including sodium oxide (Na2O), calcium oxide (CaO) or copper (II) oxide (CuO)

- Mainly Group I and II oxides (with transition metals)
- (i) $CaO + 2HCl \rightarrow CaCl_2 + H_2O$ (where the Ca^{2+} ions replace the H^+ ions)

- Amphoteric Oxides: react with both acids and bases to form salts
    - Only the "ZAP" (Zinc, Aluminium, Lead) oxides are amphoteric
    - Including zinc oxide (ZnO), Aluminium oxide (Al_2O_3), Lead (II) oxide (PbO)
    - Example: $Al_2O_3+6HCl \rightarrow 2AlCl_3+3H_2O$

Non-Metal Oxides:

- Acidic Oxides: react with bases to form salts
    - Including carbon dioxide (CO_2), sulfur dioxide (SO_2)
    - Many of them are major air pollutants
    - Usually non-metals (that react with bases)
    - $CO_2+Ca(OH)_2 \rightarrow CaCO_3+H_2O$
- Neutral Oxides: do not react with acids or bases

<u>Solubility Table:</u>

- Rule 1: Most chlorides, bromides and iodides are soluble in water. The two common exceptions are silver and lead (II) chloride/bromide/iodide
- Rule 2: Most sulfates are soluble. The three common exceptions are barium sulfate, calcium sulfate and lead (II) sulfate
- Rule 3: All carbonates are insoluble in water except ammonium carbonate and group I carbonates
- Rule 4: All oxides and hydroxides are insoluble in water except Group (I) oxides/hydroxides. (Calcium oxide and hydroxide are sparingly soluble)
- Rule 5: All group I and ammonium compounds are soluble in water
- Rule 6: All nitrates are soluble in water

<u>Precipitation Reactions:</u>

- Precipitation is the formation of a solid when 2 aqueous solutions are mixed. An insoluble substance would appear as a precipitate when it is formed.
- To form the precipitate, one of the reactants used must contain the cation (positive ion) of the ppt and the other reaction must contain the anion (negative ion) of the ppt. Both ions must be in aqueous solutions in the reactants
- A precipitate must be formed by mixing 2 aqueous solutions. It should not be formed by mixing a solid with a solution.
    - (i) $PbO\ (s) + H_2SO_4\ (aq) \rightarrow PbSO_4\ (s) + H_2O\ (l)$
    - This does not work since the solid $PbSO_4$ would form as a layer on the solid PbO and cover it, thus stopping the reaction prematurely as the acid could no longer react with the remaining PbO that is covered by the $PbSO_4$.
- Note: Aqueous lead (II) nitrate and dilute hydrochloric acid
    - The insoluble salt formed will be lead (II) chloride
    - *In this case, hydrochloric acid does not act as an acid, but as a salt

    - There is no acid-salt reaction, so the salt reacts with another salt to form an insoluble precipitate and nitric acid
- Formula of insoluble precipitates
    - For original solution containing cation (of ppt.), use nitrate anion
    - For original solution containing anion (of ppt.), use Grp I cation

<u>Ionic Reactions:</u>

- An ionic reaction specifically shows which atoms/ions/molecules took part in the reaction, through eliminating the spectator ions
- Spectator ions are free ions present at the start and at the end of a reaction (they remain unchanged by the reaction)
- The basic steps:
    1) Write the balanced equation ($Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3$)
    2) Put in the <u>state symbols</u> ($Pb(NO_3)_2$ (aq) +2KI (aq) $\rightarrow$ PbI_2 (s) + $2KNO_3$ (aq))
        - Water-soluble substances are in the aqueous state
    3) Write out the free ions in the aqueous solution
       Pb^{2+} (aq) + $2NO_3^-$ (aq) + $2K^+$ (aq) +$2I^-$ (aq) $\rightarrow$ PbI_2 (s) + $2K^+$ (aq) +$2NO_3^-$ (aq)
        - <u>Substances in solid, liquid or gaseous state do not form free ions</u>
    4) Cancel away the corresponding spectator ions on the left hand side and right hand side of the equation ($2K^+$, $2NO_3^-$)
    5) Write down the uncancelled formula gives the ionic equation
       Pb^{2+} (aq) + $2I^-$ (aq) $\rightarrow$ PbI_2 (s)
- The number of each type of atom on both sides of the equation are cancelled, the net charge on both sides are also balanced
- All neutralization reactions have the same ionic reactions (the soluble salt formed in the reaction remains as ions in the solution, while water becomes a liquid)
  <u>H^+ (aq) + OH^- (aq) $\rightarrow$ H_2O (l)</u>
- Acid metal reaction: the metal usually displaces the H^+ ion of the acid
  Zn (s) + H_2SO_4 (aq) $\rightarrow$ $ZnSO_4$ (aq) +H_2 (g)
  <u>Zn (s) + $2H^+$ (aq) $\rightarrow$ Zn^{2+} (aq) + H_2 (g)</u>
- Acid and a soluble carbonate (the hydrogen ion and the carbonate form water and carbon dioxide)
  $2HNO_3$ (aq) +Na_2CO_3 (aq) $\rightarrow$ $2Na_2NO_3$ (aq) + H_2O (l) + CO_2 (g)
  <u>$2H^+$ (aq) + CO_3^{2-} (aq) $\rightarrow$ H_2O (l) + CO_2 (g)</u>
- Displacement reaction (the more reactive metal gains electrons from the less reactive metal)
  Zn (s) + $CuSO_4$ (aq) $\rightarrow$ $ZnSO_4$ (aq) + Cu (s)
  <u>Zn (s) + Cu^{2+} (aq) $\rightarrow$ Zn^{2+} (aq) + Cu (s)</u>