- Topic 8 - Acids and Bases

$pH = -\log_{10}[H^+]$?

- Essential ideas
  - Many reactions involve the transfer of proton from an acid to a base.
  - Characterization of an acid depends on empirical evidence such as the production of gases in reactions with metals, color change of indicators, or release of heat in reactions with metal oxides and hydroxides.
  - pH scale is used to distinguish between acids, bases/alkaline, and neutral substances.
  - pH depends on concentration of solution.
    - The strength of acids and bases depends on the extent to which they dissociate (split into separate smaller atoms, ions, or molecules) in aqueous solution.
- 8.1 Theories of acids and bases - Many reactions involve the transfer of a proton from an acid to a base
  - Bronsted - Lowry theory
    - The Brønsted-Lowry Theory of acids & bases involves the transfer of protons or hydrogen ions (not an electron) within an aqueous solution. Does acidity of a solution depend on the H^+ ions in substance?
    - An acid is defined as a molecule or ion that acts as a proton donor.
      - A molecule is an electrically neutral group of two or more atoms held together by a chemical bonds.
    - A base is defined as a molecule or ion that acts as a proton acceptor.
    - E.g $HCl_{(g)}$ is dissolved in water it reacts to form H_3O^+.
      - $HCl_{(g)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$
        - In the forward reaction the hydrogen chloride acts as an acid because it donates a proton or H^+ ions, while the H_2O molecule acts as a base since it accepts the proton/ H^+ ion to form a oxonium or hydronium ion, $H_3O^+_{(aq)}$.
          - An Oxonium ion is any oxygen cation with three bonds (H_3O^+ is the simplest). ← Explain what oxonium is...
          - Hydronium is the common name of the aqueous cation H_3O^+.
        - For the reverse reaction, the acid is $H_3O^+_{(aq)}$ as it donates its excess proton to the chlorine to form HCl. The chlorine atom acts as a base as it accepts the proton.
        - The equation can be split into two "half-equations" which clearly show the proton transfer.
          - $\underset{\text{acid}}{HCl_{(aq)}} \rightleftharpoons \underset{\text{conjugated base}}{Cl^-_{(aq)}} + H^+_{(aq)}$
        - The reaction shows that when a species loses a proton (is an acid), the product has to be a base since the process is reversible (depending on the acid).
        - The chloride ion is described as the conjugate base of the hydrogen chloride molecule.
          - A conjugated base is what's left over after an acid has donated a proton (H^+).
          - $\underset{\text{base}}{H_2O_{(l)}} + H^+_{(aq)} \rightleftharpoons \underset{\text{conjugated acid}}{H_3O^+_{(aq)}}$
        - The reaction shows that when a species gains a proton (Base), its product is going to be an acid since process is reversible.
        - The hydronium or oxonium ions is described as the conjugated acid of water molecule.
          - A conjugated acid is a chemical compound formed by the reception of a proton (H^+) by a base.
        - An acid-base reaction always involves at least two conjugated pairs that differ by H^+.
        - $\underset{\text{base}}{NH_{3(g)}} + \underset{\text{acid}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{conjugate acid}}{NH_4^+{}_{(aq)}} + \underset{\text{conjugate base}}{HO^-_{(aq)}}$ How do we determine if a substance is acidic?
          - Electronegativity, the more electronegative, the stronger the acid. Bond between H^+ ion and ion X^-.
            - Although, acidity is also affected by size of atoms in bond. The larger the bonded atoms, the more acidic as the bond gets weaker, the acid becomes stronger.
          - The ammonia is acting as a base by accepting a proton from the water.
            - Water is acting as an acid as it's donating a proton (H^+ ion). When it is reacted with an acid it'll act as a base.
            - H_2O is amphiprotic, it able to act as both an acid or base depending on species reacting with.
  - Lewis acids and bases
    - A Lewis acid is a chemical species that contains an empty orbital which is capable of accepting an electron pair. Lewis acid = Electron acceptor
    - A Lewis base is a chemical species which has a filled orbital containing an electron pair which is not involved in bonding (lone pair) but may form a dative bond with a Lewis acid to form a Lewis adduct. Lewis base = Electron donator
      - A Lewis acid-base adduct is a molecule formed by the bonding of a Lewis acid with a Lewis base, without simultaneous loss of a leaving group.
        - A Lewis acid - Lewis base reaction can't be a substitution reaction.
    - Deducing the Brønsted-Lowry acid & base in a chemical reaction
      - Chloric (VII) acid, $HClO_4$, acts a monoprotic acid in water. Write an equation showing its dissociation or ionisation in water.
        - $\underset{\text{(acid)}}{HClO_{4(aq)}} + \underset{\text{(base)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{(conjugate acid)}}{H_3O^+_{(aq)}} + \underset{\text{(conjugate base)}}{ClO_4^-{}_{(aq)}}$
        - $HClO_{4(aq)}$ is an acid meaning it will donate a proton (H^+ ion), it goes on to form $ClO_4^-{}_{(aq)}$.

- $H_2O_{(l)}$ will act as a base, meaning it will accept the H^+ that the acid is releasing. It forms $H_3O^+_{(aq)}$.
- Phenylamine, $C_6H_5-NH_2$, is monoacidic. ← Ask what this is!!
  - It reacts with only one molecule of a monobasic (monoprotic acid).
  - $\underset{\text{(Base)}}{C_6H_5-NH_{2(aq)}} + \underset{\text{(acid)}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{conjugate base}}{OH^-_{(aq)}} + \underset{\text{conjugate acid}}{C_6H_5NH_3^+{}_{(aq)}}$
- The $C_6H_5-NH_2$ is a base as its a proton acceptor, while H_2O acts as an acid as its a proton donator.

- Questions
  - 1) a) NaH = acid, H_2O = base
    - b) H_2O = Base, HCO_3 = acid
    - c) H_2O = Acid, H_2NCONH_2 = base
    - d) H_3O^+ = Acid, HSO_4^- = Base
    - e) NH_4^+ = Acid, OH^- = Base

- Amphiprotic species
  - Amphiprotic means the substances can donate & accept H^+ ions.
    - Amphoteric is a general term meaning it can react both as an acid and a base.
  - H_2O can be considered both amphoteric and amphiprotic:
    - $H_2O_{(l)} + H^+_{(aq)} \rightleftharpoons H_3O^+_{(aq)}$ water acting as base
    - $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$ water acting as acid
  - All amphiprotic substances are amphoteric, because they can donate a proton when acting as an acid, and accept it when acting as a base.
  - Glycine, H_2N-CH_2-COOH has two functional groups: Explain this
    - The amine group, $-NH_2$, is a basic due to the presence of a lone pair of electrons on nitrogen atom.
    - Carboxylic acid group is acidic due to the presence of an acidic or ionisable hydrogen atom.
    - In solution and in the solid state there is an internal acid-base transfer of a proton from the carboxylic acid group to the amine group.
      - A dipolar ion or zwitterion is formed: $H_3N^+-CH_2-COO^-$.

Explain
  - Lewis acids and bases either accept or donate electron lone pairs.
    - E.g metal oxide $Mg(OH)_2$, when placed in H_2O can dissociate to release hydroxide ions:
      - $Mg(OH)_{2(aq)} \rightleftharpoons Mg^{2+}_{(aq)} + 2OH^-_{(aq)}$
    - This partial dissociation or ionization makes it a Brønsted-Lowry base, but Mg ion (Mg^{2+}), can also have water molecules coordinate to it with a lone pair, via dative bond formation making it Lewis acid (electron pair acceptor).

- Conjugate acid-base pairs
  - A conjugate acid is the molecule or ion formed when a proton is added to a base.
  - A conjugate base is the species formed when a proton is removed from an acid.
  - ? > A pair of species differing by a single proton is called a conjugate acid-base pair.
- Utilisation: Weak bases
  - Substances that remove excess acid via neutralisation are antacids.

- 8.2 Properties of acids and bases
  - Properties of acids and bases
    - Acids
      - Common acids are ethanoic acid, $CH_3COOH_{(aq)}$, sulfuric acid, $H_2SO_{4(aq)}$, hydrochloric acid, $HCl_{(aq)}$, and nitric acid, $HNO_{3(aq)}$.
    - pH
      - Acids have a pH value less than 7 and turn the indicator blue litmus paper red.
      - The pH value is a measure of acidity of the solution, and indicators are dyes that change color according to the pH of the solution.
    - Conductivity
      - Acids are electrolytes, meaning they undergo chemical decomposition when an electric current is passed through their aqueous solutions.
    - Reaction with metals
      - Most dilute acids react to give hydrogen gas and a solution of a salt when a reactive metal such as magnesium, iron, or zinc is added.
        - E.g:
          - $Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$
          - $Mg_{(s)} + 2HNO_{3(aq)} \longrightarrow Mg(NO_3)_{2(aq)} + H_{2(g)}$
          - $Mg_{(s)} + H_2SO_{4(aq)} \longrightarrow MgSO_{4(aq)} + H_{2(aq)}$

- $Mg_{(s)} + 2CH_3COOH_{(aq)} \rightarrow (CH_3COO)_2Mg_{(aq)} + H_{2(g)}$
- Or, ionically: $Mg_{(s)} + 2H^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$

- In general: Reactive metal + dilute acid → salt + hydrogen
- As you know, reactive metals are metals which don't have 2, 8, 18 etc electrons. The most reactive being the elements which are 1-2 electrons away from a full shell.
- The more unreactive metals don't react with dilute acids.

- Reaction with metal carbonates
  - Dilute acids react to give carbon dioxide gas when a metal carbonate or metal hydrogencarbonate is added.
    - E.g:
      - $\overbrace{CaCO_3}^{\text{Alkaline earth metal (Ca)}}{}_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$
      - $\underbrace{NaHCO_3}_{\text{Alkali metal (Na)}}{}_{(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$
    - Ionically:
      - $CO_3^{2-}{}_{(aq)} + 2H^+_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$
      - $HCO_3^-{}_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$
  - In general: metal carbonate or metal hydrogencarbonate + dilute acid → salt + water + carbon dioxide
  - The reaction between calcium and dilute sulfuric acid is slow because an almost insoluble layer of calcium sulfate, $CaSO_4$, protects the calcium carbonate from further attack by the acid.
  - The presence of carbon dioxide can be confirmed by bubbling the gas through limewater (a solution of calcium hydroxide).
    - Initially turns cloudy, but then clears if excess CO_2 is passed through the limestone.
      - $CO_{2(g)} + Ca(OH)_{2(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$
      - $CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \rightarrow Ca(HCO_3)_{2(aq)}$

- Reaction with bases
  - Bases include metal oxides, metal hydroxides, and aqueous ammonia.
  - A base is a substance that reacts with an acid to form a salt and water only.
    - This reaction is known as neutralization.
  - Alkalis are bases which are soluble in water.
    - These include group I hydroxides, barium hydroxide, aqueous ammonia, $NH_{3(aq)}$, sometimes called "ammonium hydroxide", $NH_4OH_{(aq)}$.
  - Alkalis have a soapy feel, and have a bitter taste.

- Reaction with metal oxides
  - Dilute acids react to give a salt and water when a metal oxide is added.
    - E.g
      - $\overbrace{CuO}^{\text{metal oxide}}{}_{(s)} + \overbrace{H_2SO_4}^{\text{Dilute acid}}{}_{(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$
      - $CuO_{(s)} + 2HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + H_2O_{(l)}$
      - $CuO_{(s)} + 2HCl_{(aq)} \rightarrow CuCl_{2(aq)} + H_2O_{(l)}$
      - $CuO_{(s)} + 2CH_3COOH_{(aq)} \rightarrow Cu(CH_3COO)_{2(aq)} + H_2O_{(l)}$
    - Ionically:
      - $O^{2-}{}_{(s)} + 2H^+_{(aq)} \rightarrow H_2O_{(l)}$
  - In general: metal oxide + dilute acid → salt + water

- Reaction with metal hydroxides
  - Dilute acids react to give a salt and water when a metal hydroxide or aqueous ammonia is added.
  - In general: metal hydroxide + dilute acid → salt + water
  - E.g:
    - $\underbrace{Na}_{\text{metal}}\overbrace{OH}^{\text{hydroxide}}{}_{(aq)} + HNO_{3(aq)} \rightarrow NaNO_{3(aq)} + H_2O_{(l)}$
    - $NH_{3(aq)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}$

- Salts
  - A salt is an ionic compound formed when the replacable hydrogen of an acid is completely or partly replaced by a metal ion.
    - Salts are an ionic compound that result from the neutralization reaction between an acid or a base.
  - Salts are formed by exothermic neutralization reactions.
  - Example:
    - $\overbrace{H}^{\text{Hydrogen (acid)}}Cl_{(aq)} \rightarrow \overbrace{Na}^{\text{metal atom}}Cl_{(aq)}$ (salt)
  - The number of replacable hydrogen atoms in an acid is called the basicity or proticity of the acid. 1 H^+ atom proticity / basicity = 1, 2 H^+ atoms = 2, etc.

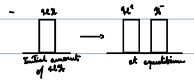
- In the case of a diprotic or dibasic, containing more than one replacable hydrogen atom, salts can be formed when all or some of the hydrogen are replaced.
    - salts formed when all hydrogen atoms are normal salts, and salts where only some of the hydrogens are replaced by metal ions they're known as acid salts.
- Aqueous solutions of salts may be neutral, acidic, or basic.

- <u>Balanced chemical equations for the reactions of acids</u>
    - $H_2SO_4 + CuCO_3 \longrightarrow H_2O_{(l)} + CO_{2\,(g)} + CuSO_{4\,(aq)}$
    - $2HBr + Ca(HCO_3)_2 \longrightarrow 2H_2O_{(l)} + 2CO_{2\,(g)} + CaBr_{2\,(aq)}$
    - $3Na_2CO_{3\,(s)} + 2H_3PO_4 \longrightarrow 3H_2O_{(l)} + 3CO_{2\,(g)} + 2Na_3PO_4$
    - $Ca_{(s)} + 2CH_3COOH_{(aq)} \longrightarrow Ca(CH_3COO)_2 + H_2$
    - $2(NH_4)_3PO_{4\,(aq)} + 3Ba(OH)_{2\,(aq)} \longrightarrow 3H_2O_{(l)} + Ba(PO_4)_{2\,(aq)} + 6NH_{3\,(g)}$

- <u>The importance of water</u>
    - Pure or anhydrous acid don't behave as acids and don't exhibit the characteristics properties of acids described previously.
        - These properties are only shown after the acids have been reacted, and dissolved in water to form an aqueous solution.
    - when placing the chemicals in water, ions will be formed: $HCl_{(g)} + (aq) \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$

- <u>8.3 The pH scale</u>
    - <u>The pH scale</u>
        - pH scale used to determine the acidity, alkalinity, or neutrality of an aqueous solution
            - pH below 7 is acidic
            - pH equal to 7 is neutral
            - pH more than 7 is basic/alkaline
        - The lower the pH the more acidic the solution.
            - Vice versa as well.
        - The pH scale is logarithmic to the base of 10, meaning a change in one unit in the pH scale will mean a change in in the hydrogen ions by an order of 10.
            - Aqueous solution with pH value of 4 is 10 times more acidic than an aqeous value of 5, and 100 times more acidic than aqueous solution of 6.
        - pH 8 is 100 times less basic than pH 10.
        - pH is directly related to the concentration of hydrogen ions present in solution.
            - pH value x corresponds to hydrogen ion concentration 10^{-x} mol dm^{-3}.
        - Hydrogen ions are present in neutral and alkaline aqueous solution because water itself is very slightly dissociated into hydrogen and hydroxide ions:
            - $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + HO^-_{(aq)}$
        - $[H^+_{(aq)}] = [OH^-_{(aq)}]$ is neutral, $[H^+_{(aq)}] > [OH^-_{(aq)}]$ is acidic, $[H^+_{(aq)}] < [OH^-_{(aq)}]$ is basic.
        - pH is the negative of the logarithm (to the base 10) of the concentration of hydrogen (or oxonium) ions.
            - pH 4 = 10^{-4} H^+ ions mol dm^{-3} = 0.0001 mol dm^{-3}, pH 11 = 10^{-11} H^+ ions mol dm^{-3}.
            - $pH = -\log_{10}[H^+]$
    - <u>The ionic product constant of water</u>
        - If the equilibrium law is applied to $H^+_{(aq)} + OH^-_{(aq)}$ we get:
            - $k_w = [H^+_{(aq)}][OH^-_{(aq)}]$
            - This is ionic product constant of water (k_w).
        - The concentration in pure water of $H^+_{(aq)}$ and $OH^-_{(aq)}$ are $1 \cdot 10^{-7}$ mol dm^{-3}, Therefore:
            - $k_w = [H^+_{(aq)}][OH^-_{(aq)}]$
              $= [1 \cdot 10^{-7}][1 \cdot 10^{-7}]$
              $k_w = 1 \cdot 10^{-14}$ mol^2 dm^{-6}
        - $[H^+_{(aq)}]$ and $[OH^-_{(aq)}]$ is a constant at a given temp.
            - As concentration of H^+ increases, OH^- concentration decreases and vice versa.
        - <u>Questions</u>
            - $pH = -\log[0.1]$
              $pH = 1$
            - 3.646 g HCl, 250 cm^3 mol $= \frac{3.646}{36.46094}$ $= 0.1$ mol $c = \frac{0.1}{0.25}$ $= 0.04$ mol dm^{-3}

- <u>8.4 Strong and weak acids and bases</u>
    - <u>Strong and weak acids</u>
        - A strong acid is an acid that is completely dissociated or ionized in an aqueous solution.
            - when a strong acid dissolves, virtually all the acid molecules react with the water to produce hydrogen (H^+) or oxium ions (OH^-).

- In general for a strong acid, HX:
  - $\underset{0\%}{HX} \rightarrow \underset{100\%}{H^+_{(aq)}} + X^-_{(aq)}$ or $\underset{0\%}{HX} + H_2O_{(l)} \rightarrow \underset{100\%}{H_3O^+_{(aq)}} + X^-_{(aq)}$
- Examples
  - $HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$
  - $HNO_{3(l)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO_3^-_{(aq)}$
  - $H_2SO_{4(l)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HSO_4^-_{(aq)}$
  - $HClO_{4(l)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + ClO_4^-_{(aq)}$
- In a strong acid:
  - 

- Monoprotic organic acids are usually weak.
- When a weak acid dissolves in water, only a small percentage (Typically 1%) react with water to release hydrogen or oxonium ions.
  - An equilibrium is established, with the majority of the acid molecules not undergoing ionisation or dissociation.
- In other words, the equilibrium lies on the left-hand side of the equation.
- In general for a weak acid, HX:
  - $HX \rightleftharpoons H^+_{(aq)} + X^-_{(aq)}$ or $\underset{99\%}{HX + H_2O_{(l)}} \rightleftharpoons \underset{1\%}{H_3O^+_{(aq)}} + X^-_{(aq)}$

- E.g of weak acids:
  - $CH_3COOH_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$
  - $CO_{2(g)} + 2H_2O_{(l)} \rightleftharpoons HCO_3^-_{(aq)} + H_3O^+_{(aq)}$

- Strong and weak bases
  - A strong base undergoes almost 100% ionisation or dissociation when in a dilute aqueous solution.
  - Strong bases have high pH values and high conductivities.
  - Comparison of weak to strong acids:

	0.1 mol dm^{-3} NaOH(aq)	0.1 mol dm^{-3} NH$_3$ (aq)
[OH$^-$(aq)]	0.1 mol dm^{-3}	≈ 0.0013 mol dm^{-3}
pH	13	11-12
Electrical conductivity	high	low

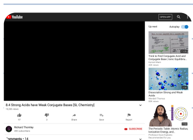
  - Like strong acids, strong bases, BOH:
    - $\underset{0\%}{BOH} + (aq) \rightarrow B^+_{(aq)} + \underset{100\%}{OH^-_{(aq)}}$

  - Three common strong bases
    - $NaOH + (aq) \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
    - $KOH + (aq) \rightarrow K^+_{(aq)} + OH^-_{(aq)}$
    - $Ba(OH)_2 + (aq) \rightarrow Ba^{+}_{(aq)} + 2OH^-_{(aq)}$
  - All bases are weak except the hydroxides of group 1 and 2.
  - Weak bases are composed of molecules that react with water molecules to release hydroxide ions ($OH^-_{(aq)}$).
  - For weak molecular base (general rule), BOH:
    - $\underset{99\%}{BOH} + (aq) \rightleftharpoons B^+_{(aq)} + \underset{1\%}{OH^-_{(aq)}}$

  - An equilibrium is established, with the majority of the base molecules not undergoing ionisation or dissociation.
    - Equilibrium lies on the left side of the equation.
  - Weak bases have low pH values, and low conductivity values.
    - Eg:
      - $NH_{3(g)} + (aq) \rightleftharpoons NH_4^+_{(aq)} + OH^-_{(aq)}$
      - $C_2H_5NH_{2(g)} + (aq) \rightleftharpoons OH^-_{(aq)} + C_2H_5NH_3^+_{(aq)}$
  - Properties of strong and weak acids
    - Strong and weak acids differ in:
      - A weak acid has a lower concentration of H$^+$ and therefore a higher pH than a strong acid of the same concentration.

- A weak acid, because of its lower concentration of H^+ ions will be a worse electrical conductor.
- Weak acids react slowly with reactive metals, metal oxides, metal oxides, metal carbonates, and metal hydrocarbonates than strong acids of the same concentration.
  - Due to low H^+.
- Strong & weak acids have different enthalpies.
- Acid strength doesn't change as the acid is diluted (at constant temp).

- Acids and their conjugates
  - $\underset{\text{acid}}{HCl_{(g)}} + \underset{\text{base}}{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$
    conjugate acid
    conjugate base
  - $\underset{\text{acid}}{CH_3COOH_{(l)}} + \underset{\text{base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{acid}}{H_3O^+} + \underset{\text{base}}{CH_3COO^-}$
    conjugate acid
    conjugate base


Video explaining why Strong acids have weak conjugate bases, and why strong bases have weak conjugate acids.

  - In the case of HCl the water molecule is a much stronger base than the chlorine ion.
    - This means that the water molecule has a greater tendency to accept a proton, $H^+_{(aq)}$, than the Cl ion.
  - The position of the equilibrium is to the right as virtually all of the hydrogen chloride molecules will ionise or dissociate.
  - In general, strong acids produce relatively weak conjugate bases in aqueous solutions.
    - The reason that they produce weak conjugate bases is because if a strong acid would form a strong base that would mean that the reverse reaction would be formed just as quickly, reassociating the dissociated acid.
  - In general, weak acids produce relatively strong conjugate bases in aqueous solutions.
  - Strong bases form weak acids, weak bases form strong acids.
  - Acids which donate a single H^+ ion are called monoprotic. Two protons = diprotic.
  - For a substance to be an acid the hydrogen has to be attracted to oxygen or a halogen.
  - In a reaction with two acids, the weaker acid will be forced to act as the base:
    - $\underset{\text{base}}{HNO_{3(aq)}} + \underset{\text{acid}}{H_2SO_4} \rightleftharpoons \underset{\text{acid}}{H_2NO_3^+{}_{(aq)}} + \underset{\text{base}}{HSO_4^-{}_{(aq)}}$
      conjugate acid
      conjugate acid

- Bases and their conjugates
  - For the weak bases of ammonia and carbonate ions the competition is between the base and its conjugate for a proton, H^+:
    - $\underset{\text{base}}{NH_{3(g)}} + \underset{\text{acid}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{acid}}{NH_4^+{}_{(aq)}} + \underset{\text{base}}{OH^-_{(aq)}}$
      99% 1%
    - $\underset{\text{base}}{CO_3^{2-}{}_{(aq)}} + \underset{\text{acid}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{acid}}{HCO_3^-{}_{(aq)}} + \underset{\text{base}}{OH^-_{(aq)}}$
  - In the case of ammonia solution, the hydroxide ion is a much stronger base than the ammonia molecule.
    - The OH^- ion has a much greater tendency to accept a proton than ammonia. Meaning position of equilibrium is to the left & few of NH_3 will dissociate in the water.
  - Weak bases produce strong conjugate bases in aqueous solutions. Strong bases form weak conjugate bases.

- 8.5 Acid deposition
  - Pollution refers to the change in equilibrium of biological & non-biological systems.
  - Primary pollutants are substances which are emitted directly from the source and remain unchanged once they enter the environment.
  - Secondary pollutants are formed in the atmosphere by chemical reactions involving primary pollutants and gases normally present in the air.
  - Formation of acid rain
    - $S + O_2 \rightarrow SO_2 + H_2O \rightarrow H_2SO_3$
      In coal or fuels oxidise during combustion
    - Sulfur dioxide undergoes photochemical oxidation in the atmosphere.
      - Occurs in water droplets.
      - Catalysed by soot or metallic particles.
      - $2SO_2 + O_2 \rightarrow 2SO_3 + H_2O \rightarrow H_2SO_4$
  - Environmental effects of acid deposition
    - Affects pH of lakes and rivers, impacting life there
    - Affects availability of metal ions in soil, affecting nearby plant life.
    - Directly affects plants.

- Affects buildings and other materials
- Directly affects human health

- **Measures to counteract acid deposition**
  - Nitrogen oxides are removed from vehicle emissions using catalytic converters.
  - Limestone increases pH in liquids & soil