

ACIDS & BASES - 1. - Chapter 15

COURSE OBJECTIVES

4. Acids and bases in aqueous solutions

- 4.1 Classify water as a weak electrolyte in terms of its partial ionisation into $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$.
- 4.2 State the expression for the ionisation constant for water, $K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$, which equals $1.0 \times 10^{-14} \text{ mol.L}^{-1}$ at 298 K.
- 4.3 State that the concentrations of $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ in pure water are equal to $1.0 \times 10^{-7} \text{ mol L}^{-1}$ at 298 K.
- 4.4 Use the Arrhenius and Bronsted-Lowry acid-base models to describe acid-base behaviour.
- 4.5 Identify in chemical equations the reactants which are acting as acids or bases.
- 4.6 Write equations to represent:
 - a) acids as sources of $\text{H}^+(\text{aq})$ and bases as sources of $\text{OH}^-(\text{aq})$ in aqueous solutions
 - b) acids as proton donors and bases as proton acceptors.
- 4.7 Define amphoteric metals, oxides and hydroxides as those which react with both acids and bases.
- 4.8 Identify chromium, zinc and aluminium as amphoteric metals forming amphoteric oxides and hydroxides and write equations for the reactions of the oxides and hydroxides of chromium (III), zinc and aluminium with $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$.
- 4.9 State and apply the relationship between pH and acidity and alkalinity in aqueous solutions.
- 4.10 Define 'pH' as $-\log_{10}[\text{H}^+(\text{aq})]$ and calculate the pH of strong acid solutions and strong base solutions.
- 4.11 Calculate the $[\text{H}^+(\text{aq})]$, $[\text{OH}^-(\text{aq})]$ and pH when solutions are mixed.
- 4.12 Identify strong acids including HCl , H_2SO_4 , HNO_3 and strong bases including NaOH , KOH , $\text{Ca}(\text{OH})_2$.
- 4.13 Identify weak acids including H_3PO_4 , CH_3COOH and other organic acids, and weak bases including NH_3 , Na_2CO_3 .
- 4.14 Distinguish between the following terms when referring to acids and bases:
 - a) 'strong' and 'weak'
 - b) 'concentrated' and 'dilute'
- 4.15 Define 'neutralisation' as a reaction of an acid with a base.
- 4.16 Describe the formation of salts by neutralisation reactions between acids and bases.
- 4.17 Use the concept of hydrolysis to explain and predict the acidic, basic or neutral nature of aqueous solutions of salts derived from the reaction of:
 - a) strong acids and weak bases (eg. ammonium chloride)
 - b) strong bases and weak acids (eg. sodium carbonate)
 - c) strong acids and strong bases (eg. sodium chloride).
- 4.L.1 Demonstrate the common properties of acids and bases.
- 4.L.2 Use indicators to determine hydrogen ion concentration and pH.

7. Inorganic chemistry

Periodic table

- 7.2 Describe the following trends across rows in the periodic table as illustrated by the third (sodium-argon) row:
 - d) acidic and basic properties of the oxides and the hydroxides.

RELEVANT TEXTBOOK SECTIONS AND PROBLEMS

Chapter 15.

- | | |
|---|--|
| 15.1 Properties of acids and bases | Rev Ex 15.1, p. 304 - 1, 2, 3 |
| 15.2 Theories of acids and bases | Rev Ex 15.2, p. 309 - 1, 2, 3 i) |
| 15.3 Ionisation of water | Rev Ex 15.3, p. 311 - 1, 2 |
| 15.4 The pH acidity scale | Rev Ex 15.4, p 314 - 1, 2, 3 |
| 15.5 Strengths of acids and bases | Rev Ex 15.5, p. 317 - 1, 2, 3 |
| 15.6 Acid ionisation constants (no calculations required) | Rev Ex 15.6, p. 319 - 1, 2 |
| 15.7 Polyprotic acids and bases | Rev Ex 15.7, p 321 - 1, 2, 3, 4 |
| 15.8 Acid-base neutralisation reactions | Rev Ex 15.8, p 322 - 1, 2 |
| 15.9 Salts | Rev Ex 15.9, p. 325 - 1, 2 |
| 15.10 Acids and bases in the living world | Rev Ex 15.10, p 326 - 1, 2 |
| 15.11 Acids, bases and the periodic table | Rev Ex 15.10, p. 329 - 1, 2 |
| Questions and problems p. 331 | 1, 2, 3, 4, 5, 6, 7i), 8, 9, 10, 12, 13, 14, 15, 16, 17, 18, 19, |
| 20 | |

Relevant 'STAWA Chemistry Problems' examples

pH calculations Set 15, p 70

PROPERTIES OF ACIDS AND BASES

Some common properties of aqueous solutions of acids

- turn blue litmus red
- conduct an electric current (because of 'free' ions)
- taste sour
- react with reactive metals **acid + metal → hydrogen gas + salt**
- react with carbonates and hydrogencarbonates
 - acid + carbonate → carbon dioxide + water + salt**
 - acid + hydrogencarbonate → carbon dioxide + water + salt**
- react with metal oxides and metal hydroxides
 - acid + metal oxide → water + salt**
 - acid + metal hydroxide → water + salt**
- react with metal sulfides (S^{2-}) **acid + metal sulfide → hydrogen sulfide + salt**
- react with sulfites (SO_3^{2-}) **acid + sulfite → sulfur dioxide + water + salt**

Properties of aqueous solutions of metal hydroxide bases

- turn red litmus blue
- conduct an electric current
- taste bitter
- react with acids **acid + metal hydroxide → water + salt**
- react with amphoteric metal hydroxides and oxides to form a complex ion e.g. $[Al(OH)_4]^-$, $[Cr(OH)_4]^-$ and $[Zn(OH)_4]^{2-}$
- react with amphoteric metals **metal + metal hydroxide + water → hydrogen gas + $[M(OH)_4]^{n-}$**

(**Note:** An amphoteric substance is one which reacts with an acid and a base.

Examples of amphoteric metal hydroxides and oxides are $Al(OH)_3$, $Cr(OH)_3$ and $Zn(OH)_2$, Al_2O_3 , Cr_2O_3 and ZnO

Examples of amphoteric metals are aluminium, chromium and zinc.

A complex ion consists of a central metal ion bonded to a group of negative ions or molecules (ligands).

Examples of complex ions are $[Al(OH)_4]^-$ - the tetrahydroxoaluminate ion, $[Cr(OH)_4]^-$ - the tetrahydroxochromate (III) ion, and $[Zn(OH)_4]^{2-}$ - the tetrahydroxozincate ion)

("Rules" for writing ionic equations -

- i) Write down the reactants and products taking care to represent them in the correct form
 - solutions of ionic compounds are written as 'separated' ions e.g. $Na^+(aq) + Cl^-(aq)$
 - solid ionic substances are written as the neutral formula e.g. $NaCl(s)$
 - insoluble ionic substances are written as the neutral formula e.g. $BaSO_4(s)$
 - solutions of strong acids (HCl , HNO_3 , H_2SO_4 , HBr) are written as 'separated' ions e.g. $H^+(aq) + Cl^-(aq)$
 - other molecular substances (e.g. gases, weak acids, water) in solution or pure form, are written as the neutral molecule e.g. H_2 , CO_2 , SO_2 , H_2O , H_2S
 - metals are written as the neutral atom e.g. Fe , Na

ii) Cancel out the spectator ions i.e. the ions that remain unchanged during the reaction

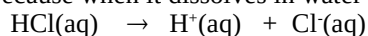
iii) Balance the equation in both atoms and charge.)

THEORIES OF ACIDS AND BASES

Arrhenius theory

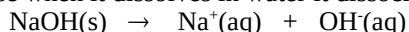
An acid is a substance that produces hydrogen ions in aqueous solution.

For example, HCl is an acid because when it dissolves in water it ionises according to the equation



A base is a substance that produces hydroxide ions in aqueous solution

For example, NaOH is a base because when it dissolves in water it dissociates according to the equation



Bronsted - Lowry theory

According to this theory, an acid-base reaction involves the transfer of a hydrogen ion (proton) from one species to another.

In an acid-base reaction, the acid is the species which donates a hydrogen ion (proton) and the base is the species that accepts the hydrogen ion (proton).

For example in the reaction $\text{HNO}_2(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{NO}_2^-(\text{aq})$

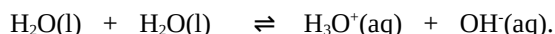
the HNO_2 is acting as the acid and the NH_3 is acting as the base.

Acid-base conjugate pair

The two species that make up an acid-base conjugate pair differ from one another by a hydrogen ion. The species with the greater number of hydrogen atoms is the conjugate acid and the other is the conjugate base. That is, the conjugate acid has one more hydrogen ion than the conjugate base. For example, Cl^- is the conjugate base of the acid HCl i.e. HCl and Cl^- are an acid/base conjugate pair.

IONISATION OF WATER

Water molecules are able to react with one another (they undergo self-ionisation) to a very small extent to produce hydronium ions and hydroxide ions:



That is, one water molecule acts as a Bronsted-Lowry acid (i.e. it donates a hydrogen ions) whilst the other acts as a Bronsted-Lowry base (i.e. it accepts a hydrogen ion).

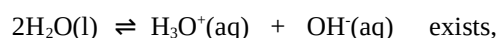
The equilibrium constant for this reaction is very small because only a few water molecules actually undergo this reaction i.e. only partial ionisation occurs. Thus, pure water is only a weak electrolyte (i.e. a poor conductor of electricity) because there are only a few ions present to conduct electricity.

The equilibrium constant for the ionisation of water (the ionisation constant of water) is given by

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ L}^{-2} \quad \text{at } 25^\circ\text{C}.$$

K_w is called the ionisation constant or the dissociation constant for water and its value depends on the temperature of the water.

In any solution in which water is the solvent i.e. in any aqueous solution, the water molecules that are present can undergo self-ionisation. Thus in any aqueous solution, the equilibrium



and at 25°C in any aqueous solution, $[\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ L}^{-2}$

$$\text{or } [\text{H}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ L}^{-2}.$$

In neutral solutions - e.g. pure water, the numbers of hydrogen ions and hydroxide ions are equal. That is, $[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ mol L}^{-1}$ (at 25°C, the pH of a neutral solution changes with temperature).

In acidic solutions - i.e. in solutions formed from acids e.g. hydrochloric acid solution, there are more hydrogen ions than in pure water i.e. the concentration of hydrogen ions is greater than 10^{-7} .

There are also hydroxide ions in an acidic solution, from the self-ionisation of water, but because in any aqueous solution $[H^+] \times [OH^-] = 10^{-14}$, the concentration of hydroxide ions must be less than 10^{-7} . That is, in acidic solutions

$[H^+] > [OH^-]$.

In basic solutions - i.e. in solutions formed from bases e.g. sodium hydroxide solution, there are more hydroxide ions than in pure water i.e. the concentration of hydroxide ions is greater than 10^{-7} .

Again, because of the equilibrium condition that exists between hydroxide ions and hydrogen ions in any aqueous solution, $[H^+] \times [OH^-] = 10^{-14}$, it follows that the concentration of hydrogen ions must be less than 10^{-7} . That is, in basic solutions $[OH^-] > [H^+]$.

THE pH SCALE

The pH of an aqueous solution is defined by $\text{pH} = -\log_{10}[H^+]$ (or $\text{pH} = -\log_{10}[H_3O^+]$)

(Note $-\log_{10} 10^{-a} = a$)

Thus, - the pH of a neutral solution is 7 because $[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ mol L}^{-1}$.

- the pH of an acidic solution is less than 7 because $[H^+] > [OH^-]$ i.e. $[H^+] > 1.0 \times 10^{-7}$

- the pH of a basic solution is greater than 7 because $[H^+] < [OH^-]$ i.e. $[H^+] < 1.0 \times 10^{-7}$

For example,

i) a $0.00100 \text{ mol L}^{-1}$ HCl solution will have a pH of 3 because the $[H^+]$ in the solution will be $1 \times 10^{-3} \text{ mol L}^{-1}$.

ii) if a solution has a pH of 4.2 then the $[H^+]$ in this solution will be $10^{-4.2}$ i.e. $6.31 \times 10^{-5} \text{ mol L}^{-1}$

Example:

10.0 g of calcium hydroxide is added to 200 mL of 0.400 mol L^{-1} HCl. Calculate the pH of the solution resulting from this reaction. Assume that the volume is not affected by the addition of the calcium hydroxide.

Solution:

moles of $\text{Ca(OH)}_2 = \frac{10.0}{74.1} = 0.1350$

moles of OH^- present before the reaction $= 2 \times 0.1350 = 0.2700$

moles of HCl $= 0.200 \times 0.400 = 0.08000 =$ moles of H^+ before the reaction

$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$

0.08000 moles of H^+ will react with 0.08000 mole of OH^- i.e. H^+ is limiting reagent.

moles of OH^- remaining after reaction $= 0.2700 - 0.08000 = 0.1900$

concentration of $\text{OH}^- = \frac{0.1900}{0.200} = 0.9500$

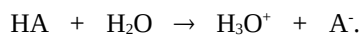
concentration of $\text{H}^+ = \frac{1.053 \times 10^{-14}}{0.9500} \text{ mol L}^{-1}$

$\text{pH} = -\log_{10}[H^+] = -\log_{10}(1.053 \times 10^{-14}) = \mathbf{14.0}$

STRENGTHS OF ACIDS AND BASES

Acids

When an acid is dissolved in water it reacts with water to form H_3O^+ (or H^+) ions. This can be represented as follows where HA is an acid

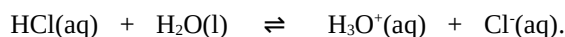


The solution formed when an acid dissolves in water is called an acidic solution because H_3O^+ is formed in the hydrolysis reaction (and so the concentration of H_3O^+ will be greater than the concentration of OH^-)

Strong Acids

Strong acids e.g. hydrochloric acid, are essentially completely ionised when placed in solution i.e. essentially all the HCl molecules donate their hydrogen ions (protons) to water molecules. That is, a solution of hydrochloric acid would contain essentially no HCl molecules, but many H^+ and Cl^- ions.

This can be explained in terms of the equilibrium system that is set up when a strong acid is added to water. For example, in hydrochloric acid, the following equilibrium system exists



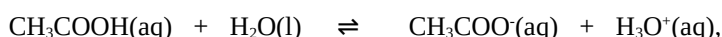
The equilibrium position greatly favours the products, so HCl is a strong acid i.e. very few HCl molecules would be present in solution; they have been ionised to form ions. The equilibrium constant for this reaction would be very large.

Five common strong acids are hydrochloric acid, sulfuric acid (H_2SO_4), nitric acid (HNO_3), perchloric acid (HClO_4) and hydrobromic acid (HBr).

Weak Acids

Weak acids e.g. acetic acid, are only partially ionised in solution, that is, only some of the dissolved acetic acid molecules donate their hydrogen ions (protons) to water molecules. Therefore, a solution of acetic acid would contain many CH_3COOH molecules and a few CH_3COO^- and H^+ ions.

For the equilibrium system that is established when acetic acid is dissolved in water i.e.



the position of equilibrium favours the reactants, so acetic acid is a weak acid i.e. the equilibrium constant for the forward reaction would be small.

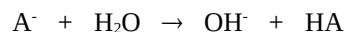
There are many weak acids, e.g. ammonium ions (NH_4^+), hydrofluoric acid (HF), phosphoric acid (H_3PO_4), sulfurous acid (H_2SO_3), carbonic acid (H_2CO_3), nitrous acid (HNO_2), hydrogensulfate ion (HSO_4^-), dihydrogenphosphate ion (H_2PO_4^-), acetic acid (CH_3COOH) and other organic acids.

(Group III and transition metal ions e.g. aluminium ions and iron(III) ions also behave as weak acids in aqueous solution. This behaviour is due to a reaction such as



Bases

When a base is dissolved in water it reacts with water to form OH^- ions. This can be represented as follows where A^- is the base



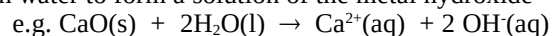
The solution formed when a base dissolves in water is called a basic solution because OH^- is formed in the hydrolysis reaction (and so the concentration of OH^- will be greater than the concentration of H_3O^+)

Strong Bases

Strong bases e.g. sodium hydroxide, completely dissociate into hydroxide ions and metal ions in aqueous solution. For example: the reaction $\text{NaOH(aq)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$ is essentially complete.

Most metal hydroxides and metal oxides are strong bases e.g. NaOH, $\text{Ca}(\text{OH})_2$, K_2O , MgO.

The metal oxides react with water to form a solution of the metal hydroxide



Weak Bases

For weak bases e.g. ammonia, only a small proportion of the molecules or ions react with water to form hydroxide ions in aqueous solution.

For example, the reaction $\text{NH}_3(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$ has only a very small equilibrium constant

There are many weak bases, e.g. ammonia, hydrogencarbonate ion, carbonate ion, CN^- , F^- , negative ions of weak acids

'Weak' and 'Strong', 'Dilute' and 'Concentrated'

The words "weak" and "strong" are used to describe the extent of dissociation/ionisation of an acid or base in aqueous solution. For example, a weak acid is less ionised in water than a strong acid.

The words "dilute" and "concentrated" are used to describe the amount of solute dissolved in a certain volume of solution. For example, a 0.1 mol L⁻¹ HCl solution would be a dilute solution, but a 10 mol L⁻¹ HCl solution would be a concentrated solution.

ACID-BASE NEUTRALISATION REACTIONS

A neutralisation reaction can be regarded as a reaction between an acid and a base.

Most acid-base neutralisation reactions result in the formation of a salt and water.

For example, when nitric acid reacts with the base sodium hydroxide, the products formed are water and a salt called sodium nitrate.

SALTS

Definition

A salt is an ionic compound containing an anion (negative ion) other than OH⁻ and O²⁻.

Salts are often formed when the hydrogen of an acid is replaced by a metal ion or an ammonium ion in a reaction. For example, NaCl and NH₄Cl are salts formed from hydrochloric acid (HCl).

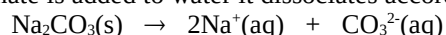
That is, the negative ion in a salt comes from the acid and the positive ion comes from the substance that reacts with the acid.

For example, when nitric acid reacts with potassium carbonate, the negative ion from the acid is the nitrate ion and the positive ion from the other substance is the potassium ion. Thus the salt formed will be potassium nitrate

Solutions of salts in water

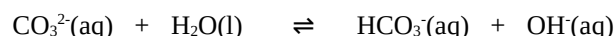
When a soluble salt is added to water, it becomes completely **dissociated** into ions (i.e. it dissolves by dissociating into 'separated' ions).

For example, when sodium carbonate is added to water it dissociates according to the following equation:



The ions that are formed in the dissociation process can then react further with water. This reaction with water is called a **hydrolysis** reaction.

For example, the hydrolysis reaction for the carbonate ion is



The products formed in the hydrolysis process determine whether the solution/salt is acidic, basic or neutral.

If H₃O⁺ ions are formed in the hydrolysis reaction then the solution, and the salt, are said to be **acidic**.

If OH⁻ ions are formed in the hydrolysis reaction then the solution, and the salt, are said to be **basic**.

If neither of the ions formed in the dissociation reaction undergo hydrolysis, then the solution, and the salt, are said to be **neutral**.

For the example given above, sodium carbonate would be called a basic salt because OH⁻ ions are formed in the hydrolysis reaction (the Na⁺ ions do not undergo hydrolysis)

Acidic/basic properties of ions.

Neutral ions

Most negative ions which are conjugate bases of strong acids e.g., Cl⁻, NO₃⁻, Br⁻, are very, very weak bases, and it is assumed that they do not hydrolyse in aqueous solution. That is, these ions are regarded as neutral ions.

Metal ions of groups I and II e.g. Li^+ , Na^+ , K^+ , Mg^{2+} , Ca^{2+} , Ba^{2+} do not react significantly with water, and are regarded as neutral ions.

Basic ions

Most negative ions which are conjugate bases of weak acids e.g. F^- , CH_3COO^- , PO_4^{3-} , HCO_3^- , CO_3^{2-} , S^{2-} , ClO^- , are weak bases. They undergo hydrolysis in aqueous solution, and produce significant amounts of hydroxide ions, thus causing the solution to be basic.

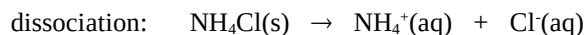
Acidic ions

A few negative ions are weak acids (in general, these are derived from polyprotic acids) e.g. HSO_4^- and H_2PO_4^- are weak acids i.e. they lose a hydrogen ion when reacting with water.

A few positive ions are weak acids e.g. Al^{3+} , Fe^{3+} , and NH_4^+ .

In summary, when a salt dissolves in water, dissociation first occurs followed by hydrolysis of the dissolved ions. If this hydrolysis is significant, then an acidic or basic solution will be obtained. If the hydrolysis is not significant, then a neutral solution is obtained.

For example, NH_4Cl is an acidic salt because it dissociates to form NH_4^+ and Cl^- ions. The Cl^- does not react further with water (hydrolyse), but the NH_4^+ does hydrolyse, to form H_3O^+ ions i.e.



(Note: - A salt formed in the reaction between a strong acid and a strong base is neutral.

- A salt formed in the reaction between a strong acid and a weak base is an acid.

- A salt formed in the reaction between a weak acid and a strong base is a base.)

ACIDS, BASES AND THE PERIODIC TABLE

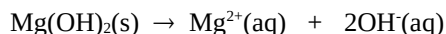
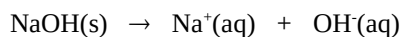
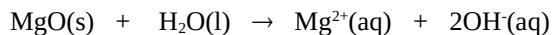
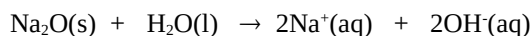
Trends across the periods

The acid-base properties of the oxide and hydroxides changes from basic, amphoteric to acidic, across a period in the periodic table.

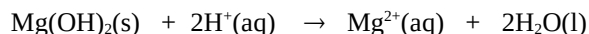
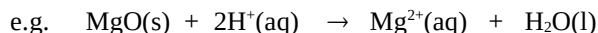
For example, for the third period

i) the oxides and hydroxides of Na, Mg are **strongly basic**.

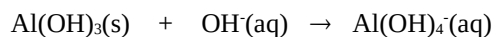
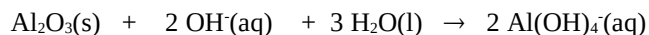
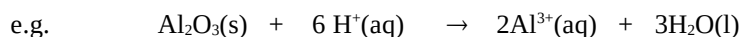
- in aqueous solution, they completely react/ dissociate to form OH^- ions



- they all react with acids, but not with bases.

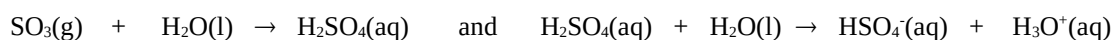


ii) the oxide and hydroxide of Al are **amphoteric** - i.e. they will react with both acids and bases

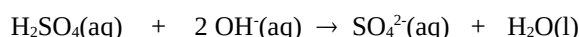
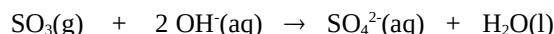
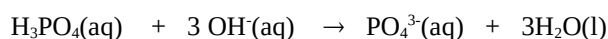
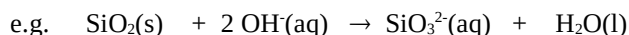


iii) the oxides and hydroxides of Si, P, S and Cl are acidic (weakly acidic → strongly acidic)

- they ionise in aqueous solution to form hydrogen ions



- they also react with strong bases, but not with acids.



Trends down the groups

Going down a group, there is a decrease in acidic properties, and an increase in basic properties.

For example, in group IV, carbon dioxide is moderately acidic, silicon dioxide and germanium dioxide are weakly acidic, and tin dioxide and lead dioxide are amphoteric.

ADDITIONAL IDEAS

STRENGTHS OF ACIDS AND BASES

The equilibrium position of an acid-base reaction can also be used to compare the strengths of acids and bases involved in the reaction.

For example, for the reaction $\text{HCl}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$, the equilibrium position is determined by which acid, the HCl or the H_3O^+ , 'wants to donate its protons the most' or alternatively, which base, the H_2O or the Cl^- , 'wants to accept hydrogen ions the most'. Because the equilibrium position is far towards the right, it means that HCl is a much stronger acid than H_3O^+ or alternatively H_2O is a stronger base than Cl^- .

Also, there is a relationship between the acid/base strengths of a conjugate pair.

For example, for the conjugate pair HCl / Cl^- , HCl is a strong acid, meaning that it very readily donates its hydrogen ion. But it also follows that its conjugate base i.e. Cl^- , must be very weak as it does not readily accept hydrogen ions to form HCl.

Thus a strong acid will have a very very weak (or neutral) conjugate base, and a strong base will have a very weak conjugate acid.

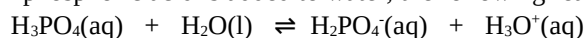
(Note - it does not necessarily follow that a weak acid will have a strong conjugate base, nor that a weak base will have a strong conjugate acid, as often a weak acid has a weak conjugate base e.g. acetic acid is a weak acid, and the ethanoate ion is a weak base.)

If the relative strengths of two acids or bases are known then the relative strengths of their conjugate pair can be determined e.g. the acid H_3PO_4 is a stronger acid than H_2S , and so, H_2PO_4^- will be a weaker base than HS^- .

ACID DISSOCIATION CONSTANTS

The equilibrium constant for the ionisation process that occurs when an acid ionises in aqueous solution is called the acid ionisation constant or the acid dissociation constant or the acidity constant (K_a).

For example, when phosphoric acid is added to water, the following reaction occurs



The acid dissociation constant for this reaction is given by the expression $K_a =$

and it has a value of 7.5×10^{-3} at 25°C .

This small value for K_a indicates that phosphoric acid is a weak acid. Strong acids have very large K_a values.