

TOPIC 13.

ACIDS AND BASES - pH calculations.

Acids and bases.

In an earlier Topic, an **ACID** was described as a species which provides H^+ ions in solution. Any species which accepts the H^+ ions is called a **BASE**. For example, the hydroxide ion is a base in the reaction $OH^- + H^+ \rightarrow H_2O$. In this Topic the concepts of acids and bases are considered in more detail.

The role of the solvent.

Consider the species HNO_3 which was observed earlier to behave as an acid. When molecules of HNO_3 are dissolved in water, they completely dissociate into ions to form aquated hydrogen ions which have previously been written as $H^+(aq)$, and nitrate ions, $NO_3^-(aq)$. For convenience, the $H^+(aq)$ ion is usually written as H_3O^+ in discussions of acids and bases, this being the simplest formula for the combination of an H^+ ion with a water molecule. This convention will be used for the remainder of the discussion here. Then the equation for the dissociation of nitric acid in water, (deleting the (aq) suffixes) is



From the equation it is seen that the solvent, water, has accepted the H^+ ion from the donor species, HNO_3 . This reaction is complete in that virtually all HNO_3 molecules react with H_2O molecules to form the ions on the right hand side, and there are no HNO_3 molecules remaining in solution. When an acid is completely dissociated in a given solvent (water in this instance) it is called a **STRONG ACID**, and there is none of the non-ionised (molecular) form remaining.

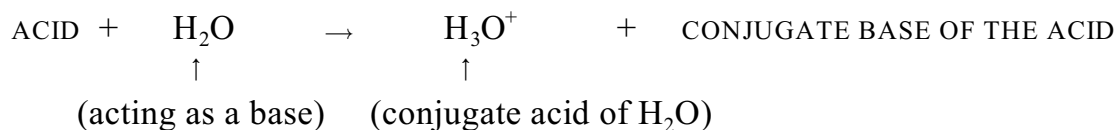
Hence it is a nonsense to write $HNO_3(aq)$ as found in many texts, because in water solution, there would be almost no undissociated molecules of HNO_3 present.

Because the solvent molecules, H_2O , accept the H^+ ions, they are acting as bases and the dissociation of HNO_3 is an acid/base reaction between the HNO_3 molecules acting as an acid and the H_2O molecules acting as the base.

Conjugate acid-base pairs.

Returning to equation (1) above, the species left after an acid has dissociated (NO_3^- in this example) is called the **CONJUGATE BASE** of the acid. A conjugate acid/base pair are any two species that differ in formula by a single H^+ . Similarly, as the difference between the formulas H_2O and H_3O^+ is a single H^+ , then this is also a conjugate acid/base pair - the H_3O^+ ion must be the acid as it has the extra H^+ in its formula, and the H_2O molecule is its conjugate base.

Thus for any acid dissociating in water, the process can be represented as



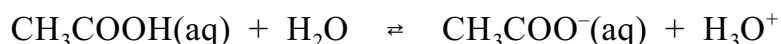
The following Table lists some commonly encountered conjugate acid/base pairs.

ACID	BASE	ACID	BASE
H_2SO_4	HSO_4^-	HSO_4^-	SO_4^{2-}
H_3PO_4	H_2PO_4^-	HI	I^-
H_2PO_4^-	HPO_4^{2-}	HBr	Br^-
HPO_4^{2-}	PO_4^{3-}	HCl	Cl^-

Note that in two of these examples the acid species was capable of donating more than a single H^+ . The sulfuric acid molecule, H_2SO_4 , has two dissociations possible and is called a **DIPROTIC ACID**, while the phosphoric acid molecule is capable of donating up to three H^+ ions and is called a **TRIPROTIC ACID**. Note also however, that each dissociation involves a particular acid/conjugate base pair such as $\text{H}_2\text{SO}_4/\text{HSO}_4^-$, and that there is no conjugate relationship between the species H_2SO_4 and SO_4^{2-} which differ in formula by **two** hydrogen ions.

Weak acids.

In the previous example (1), because the dissociation is complete, the nitric acid, HNO_3 , is a strong acid. However, there are many species which provide H^+ ions in water solution without being completely dissociated. For example, the molecule CH_3COOH (acetic acid, the main constituent of vinegar) is typical of a very large group of compounds, many of which contain the COOH group as part of their molecule and which only slightly dissociate in water. Such acids are called **WEAK ACIDS** and represent an example of a chemical **EQUILIBRIUM**. This equilibrium process is one in which reactants on the left hand side of the equation form products on the right hand side at exactly the same rate as the reverse reaction occurs. Because at any given instant some species are actually reacting yet at the same time the overall amounts of all species present is constant, the term **DYNAMIC EQUILIBRIUM** is used. In the equation for such reactions, reversible or equilibrium arrows are used instead of "one way" arrows. Thus the equation for a weak acid such as CH_3COOH dissociating in water would be



This means that CH_3COOH molecules are reacting with H_2O molecules to form their conjugates at exactly the same rate as those conjugates, CH_3COO^- ions and H_3O^+ ions, are recombining to form CH_3COOH and H_2O molecules.

Note that for water solutions of weak acids as distinct from strong acids, it is appropriate to write formulas such as $\text{CH}_3\text{COOH}(\text{aq})$ for acetic acid showing the molecular form associated with water molecules as this compound is present mostly as undissociated aquated CH_3COOH molecules. However, the (aq) suffixes are usually deleted for convenience.

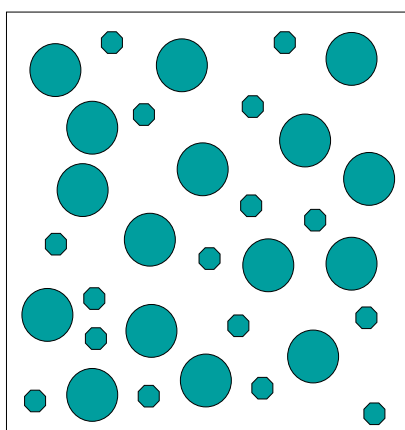
There are relatively few strong acids but many acids are weak. The following list gives some examples from both groups.

STRONG ACIDS: HCl, HBr, HI, H_2SO_4 , HNO_3 , HClO_4 , (but not HF).

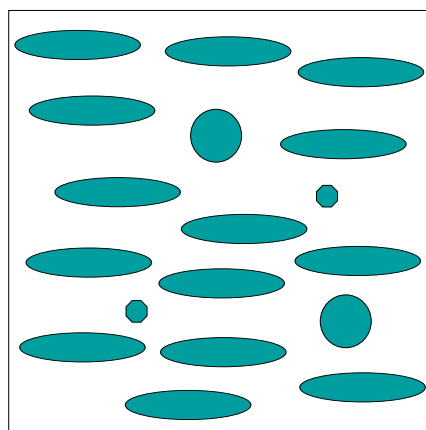
WEAK ACIDS: HF, HSO_4^- , H_3PO_4 , H_2PO_4^- , H_2CO_3 , HNO_2 , H_2SO_3 , CH_3COOH .

Strong and weak are not the same as concentrated and dilute.

In common parlance concentrated solutions are often referred to as being “strong” while dilute solutions might be called “weak”. It is important to note that in chemistry the terms “strong” and “weak” as applied to solutions have the special meanings defined on the previous pages. Strong and weak refer to the degree of dissociation of the particular species under consideration while concentrated and dilute refer to the concentration of the solute present. It is possible to have a dilute solution of a strong acid - e.g. 0.1 M nitric acid would be considered dilute while a 10 M solution of nitric acid would be deemed to be concentrated. Likewise, one can have a concentrated solution of a weak acid - e.g. 10 M acetic acid would be a concentrated solution while 0.1 M acetic acid is a dilute solution. The following diagrams illustrate the difference between a strong and a weak acid in solution. The representation on the left is of a strong acid such as HNO_3 while that on the right is of a weak acid such as CH_3COOH .



HNO_3 in solution. The small circles represent H^+ ions, the large circles represent NO_3^- ions. Dissociation into H^+ and NO_3^- ions is complete with no HNO_3 molecules remaining.



CH_3COOH in solution. The small circles represent H^+ ions, the large circles CH_3CO_2^- ions and the spherical shapes represent undissociated CH_3COOH molecules.

[Note: Acetate ions are usually represented as CH_3CO_2^- rather than CH_3COO^-]

Check your understanding of this section.

Give a definition of an acid and of a base.

What is (i) a strong acid (ii) a weak acid?

What does the term “conjugate acid base pair” mean?

Write the formula for the conjugate acid of chloride ion.

Write the formula for the conjugate base of H_2O .

Are the species H_3O^+ and OH^- a conjugate pair? Justify your answer.

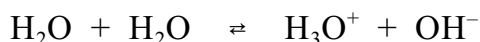
The term “dynamic equilibrium” might appear to be self-contradictory. Explain what this term means in the context of acid/base theory.

Distinguish between the terms (i) strong vs concentrated and (ii) weak vs dilute.

What is the role of water molecules in the reaction of hydrogen chloride with water?

Ionisation of water.

Electrical conductivity tests show that even the purest water has some ions present, and this is due to a very slight amount of dissociation of water molecules themselves into ions. Again, the process is a dynamic equilibrium with only an extremely small amount of the ions actually present at any instant. The process is represented by the equation



Here one of the H_2O molecules on the left hand side of the equation is acting as an acid, donating an H^+ ion to the other H_2O molecule which acts as a base to form their conjugates, H_3O^+ and OH^- on the right. Because water molecules on the left are acting as both the acid and the base, water is called a **SELF-IONISING SOLVENT**.

The amount of self-ionisation of water depends upon the temperature. At room temperature of 25°C , the observed concentration of H_3O^+ ions present in pure water is 1×10^{-7} moles per litre (M). At higher temperatures the water molecules have more energy and the equilibrium moves a little to the right, causing some increase in the concentration of H_3O^+ and OH^- ions present, while at lower temperatures less ionisation is observed. For example, at 50°C the concentration of H_3O^+ in pure water is observed to be 2.3×10^{-7} M while at 10°C it is 0.5×10^{-7} M.

The pH scale.

From the examples just cited, it can be seen that expressing the amount of H_3O^+ present in pure water or dilute solutions of acids in water involves inconveniently small numbers if moles per litre is used as the concentration unit. To overcome this inconvenience, when the amount of H_3O^+ present in a given solution is small (say less than 1 M), it is usually expressed in terms of a quantity called the **pH**.

The symbol "p" is simply shorthand for " $-\log_{10}$ " and "pH" means " $-\log_{10} [\text{H}_3\text{O}^+]$ ". The square brackets are commonly used to mean "the concentration in moles/litre of" whatever is enclosed within them. Thus, for pure water, the amount of H_3O^+ present can be conveniently given by the pH which at 25°C would be

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}(1 \times 10^{-7}) = -(-7.0) = 7.0 \text{ (no units)}$$

Similarly, the hydrogen ion concentration of dilute solutions of acids is usually expressed via the pH rather than directly in moles/litre. Acidic solutions must have a greater concentration of hydrogen ions than that present in pure water so it follows that *the pH of acidic solutions must be less than 7*, the pH of pure water.

Provided the acid is strong, the pH can be calculated as in the following examples.

Example 1. Calculate the pH of 0.10 M HNO_3 .

As the acid is completely dissociated, then the $[\text{H}_3\text{O}^+]$ in the solution is the same as that of the HNO_3 dissolved = 0.10 M.

$$\therefore \text{pH} = -\log(0.10) = -(-1.00) = 1.00$$

Example 2. Calculate the pH of 0.010 M HNO_3 .

$$\text{pH} = -\log(0.010) = -(-2.00) = 2.00$$

Notice that because of the log scale employed, a change of 1 pH unit represents a ten fold change in the concentration of hydrogen ions.

Example 3. Calculate the pH of 4.3×10^{-4} M hydrochloric acid solution.

Again the acid is strong, so the $[\text{H}_3\text{O}^+]$ is also 4.3×10^{-4} M, and

$$\therefore \text{pH} = -\log(4.3 \times 10^{-4}) = -(-3.37) = 3.37$$

Notice how the pH of acidic solutions approaches 7 from below as the acid is diluted.

Example 4. Given the pH of a solution of HNO_3 is 3.4, calculate the concentration of H^+ in that solution.

If $\text{pH} = 3.4$, then $-\log[\text{H}_3\text{O}^+] = 3.4$ and thus $[\text{H}_3\text{O}^+] = 10^{-3.4}$ M.

It is normal to express concentrations with integer powers of 10, so this would be better given as 4×10^{-4} M.

Note in this example the calculation is from pH to hydrogen ion concentration. In general terms, this can be expressed as

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} \text{ M}$$

Basic solutions.

Up to this point we have only considered water itself or acidic solutions. At 25°C , pure water has $[\text{H}^+] = 1 \times 10^{-7}$ M while acidic solutions have $[\text{H}^+]$ greater than this. In pure water, the concentrations of H^+ and OH^- ions are equal, as one of each ion results from the dissociation of a single water molecule.

$$[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7} \text{ M in pure water (at } 25^\circ\text{C}).$$

In acidic solutions, the concentration of H^+ ions exceeds that of OH^- ions - i.e. $[\text{H}^+]$ is greater than 1×10^{-7} M.

$$[\text{H}^+] > 1 \times 10^{-7} \text{ M in acidic solutions}$$

One characteristic of the equilibrium between H_2O , H^+ and OH^- is that $[\text{H}^+] \times [\text{OH}^-]$ is a constant, which at $25^\circ\text{C} = 1 \times 10^{-14} \text{ M}^2$, determined by experiment. This can be written as an equilibrium expression

$$[\text{H}^+] \times [\text{OH}^-] = 1 \times 10^{-14} \text{ M}^2 \text{ in all dilute water solutions.}$$

Thus in acidic solutions, as the $[\text{H}^+]$ is increased, then the $[\text{OH}^-]$ decreases.

Basic solutions are those where $[\text{OH}^-] > [\text{H}^+]$. Consequently, for such solutions, $[\text{H}^+]$ must be less than the value in pure water, 1×10^{-7} M. Like a see-saw, if one end goes up, the other goes down. Increased $[\text{H}^+]$ means decreased $[\text{OH}^-]$ and vice versa.

$$[\text{OH}^-] > 1 \times 10^{-7} \text{ M in basic solutions}$$

$$[\text{H}^+] < 1 \times 10^{-7} \text{ M in basic solutions}$$

The pH of a basic solution can be calculated by first using the equilibrium relation $[\text{H}^+] \times [\text{OH}^-] = 1 \times 10^{-14} \text{ M}^2$ to calculate the $[\text{H}^+]$ and then using the expression for $\text{pH} = -\log[\text{H}^+]$ as before. Note that the pH of basic solutions will always be greater than 7.

At 25°C, acidic solutions have $[H^+] > 1 \times 10^{-7} \text{ M}$ and $\text{pH} < 7$
 pure water has $[H^+] = [OH^-] = 1 \times 10^{-7} \text{ M}$ and $\text{pH} = 7$
 basic solutions have $[H^+] < 1 \times 10^{-7} \text{ M}$ and $\text{pH} > 7$

Example 5. Calculate the pH of 0.10 M sodium hydroxide solution.

The formula of sodium hydroxide is NaOH and this compound is completely soluble. Therefore $[OH^-]$ ion in solution = 0.10 M.

As $[H^+] \times [OH^-] = 1.0 \times 10^{-14}$, then $[H^+] = 1.0 \times 10^{-14} / [OH^-]$

$$\text{i.e. } [H^+] = \frac{1.0 \times 10^{-14}}{0.10} = 1.0 \times 10^{-13} \text{ M}$$

$$\text{and pH} = -\log[H^+] = -\log(1.0 \times 10^{-13}) = -(-13.00) = 13.00$$

Example 6. Calculate the pH of 0.010 M sodium hydroxide solution.

As in the previous example, because NaOH is completely soluble in water, $[OH^-]$ ion in solution = 0.010 M.

As $[H^+] \times [OH^-] = 1.0 \times 10^{-14}$, then $[H^+] = 1.0 \times 10^{-14} / [OH^-]$

$$\text{i.e. } [H^+] = \frac{1.0 \times 10^{-14}}{0.010} = 1.0 \times 10^{-12} \text{ M}$$

$$\text{and pH} = -\log[H^+] = -\log(1.0 \times 10^{-12}) = -(-12.00) = 12.00$$

Notice how the pH of basic solutions approaches 7 from above as the base is diluted.

Calculation of pH for weak acid solutions.

In the previous examples, as the acid used in each case was strong, then there was equality between the concentration of acid used and the concentration of hydrogen ions resulting. For weak acids, because they are dissociated to only a small extent, an additional step is required to connect these two quantities and will not be dealt with here. However, once the hydrogen ion concentration has been calculated using that step, the pH is calculated in the same way as for strong acids.

Indicators.

Some substances when dissolved in solution have the ability to change their colour, depending on the solution pH. These are called **ACID-BASE INDICATORS**. Indicators are themselves weak acids which are either colourless or which have a very intense colour that is a characteristic of each particular indicator. The conjugate base of the weak acid must have a different colour.

When the pH of the solution is increased to a suitable value (i.e. the solution is made less acidic or more basic), the weak acid is converted into its conjugate base by loss of an H^+ and the colour of the solution is seen to change from that of the acidic form of the indicator to the colour of its conjugate base.

There are many compounds that serve as indicators, some of them being common natural products found in plants. Each indicator has a characteristic pH at which it changes from the acid form to the base form. For example, the best known indicator, litmus, changes at $\text{pH} = 7$. Other common indicators include methyl orange which changes colour from red to yellow at $\text{pH} = 4$ and phenolphthalein which changes from colourless to red at $\text{pH} = 10$. Universal indicator contains a number of indicators in the same solution and shows the pH of solutions within the range from $\text{pH} = 4$ to $\text{pH} = 11$.

Check your understanding of this section.

What does the term self-ionising solvent mean?

What does the operator designated by the symbol p represent?

Why is pH a convenient way of expressing the concentration of dilute acid solutions?

If the pH of a solution is 1 unit greater than another solution, what would be the relative H^+ concentrations of the two solutions?

Would the pH of pure water be 7 at a temperature of 50°C ?

If a solution has a $\text{pH} = 5$, what does this mean (i) qualitatively and (ii) quantitatively regarding its acid/base properties?

What generalisation can be made regarding the pH of basic solutions?

Give the relationship which connects the concentrations of H^+ and OH^- ions in water at room temperature.

What would be the pH of a 0.10M water solution of sodium hydroxide at 25°C ?

Why can't the pH of a weak acid be calculated directly from the concentration of the acid solution?

What are the characteristics required of a useful acid/base indicator?

Objectives of this Topic.

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Know that an acid is a species which donates H^+ and a base accepts H^+ .
2. Know the meaning of the terms: strong acid; weak acid; conjugate base; conjugate acid; concentrated solution; dilute solution; dissociation.
3. Be able to write the formula for the conjugate base of any given acid or the conjugate acid of any base.
4. Know the meaning of the terms diprotic and triprotic acids.
5. Understand the concept of dynamic equilibrium as applied to weak acids, and be able to write an equilibrium equation for the partial dissociation of weak acids.
6. Know that water is slightly ionised and is capable of being both an acid and a base.
7. Know the meaning of the symbol "pH" and understand why it is used.
8. Be able to calculate the pH of a strong acid solution and be able to convert pH to concentration of H^+ .
9. Know that at $25^\circ C$, water contains H^+ ions and OH^- ions both at an equal concentration of $1 \times 10^{-7} M$ and thus has a pH of 7.
10. Know that in acidic solutions, the concentration of H^+ ions is greater than the concentration of OH^- ions and the $pH < 7$.
11. Know that in basic solutions the concentration of OH^- ions is greater than the concentration of H^+ ions and the $pH > 7$.
12. Know the relationship between the concentrations of H^+ and OH^- ions in water solutions at $25^\circ C$ by experiment is $[H^+] \times [OH^-] = 1 \times 10^{-14} M^2$ and be able to calculate the pH of a strong base in solution.
13. Know why the pH of a weak acid cannot be calculated directly from the acid concentration.
14. Understand the principles underlying the operation of acid/base indicators.

If you are unsure about the interconversion of numbers to their logarithmic form and vice versa, please consult Appendix 2 at the end of these notes or the on-line maths revision notes available from the downloads page of the School of Chemistry which also show how to use a calculator for this purpose.

www.chemlab.chem.usyd.edu.au/download.htm

SUMMARY

Acids are species which can donate hydrogen ions, H^+ , while bases are species which can accept them. A strong acid in any given solvent (usually water) is one which is completely dissociated to H^+ ions and its conjugate base. A conjugate acid/base pair is any pair of species whose formulas differ by a single H^+ ion. Some acid species can donate more than a single H^+ ion, for example H_2SO_4 and H_3PO_4 are diprotic and triprotic acids respectively.

Weak acids are those that are only partially dissociated to H^+ ions and conjugate base and remain for the most part in the non-dissociated form. Such acids exist in a dynamic equilibrium wherein acid species are dissociating at exactly the same rate as their conjugate base and H^+ ions are recombining. There are relatively few strong acids but many weak acids including organic acids containing the COOH group.

For acids that are dissociating in water as solvent, the water molecules behave as the base and accept the H^+ from the acid in an acid/base reaction. Some solvents are able to both accept and donate H^+ ions and are called self-ionizing solvents.

Water is an example of a self-ionising solvent as it can react with itself slightly to form H_3O^+ and OH^- ions. Experimentally the concentrations of each of these ions at $25^\circ C$ is found to be 1×10^{-7} M and they exist in a dynamic equilibrium with undissociated H_2O molecules. As a consequence of this equilibrium, if more H^+ is added to water by adding an acid, the concentration of H^+ in solution increases and the concentration of OH^- decreases. Alternatively, adding a base to water removes H^+ ions and so their concentration decreases while the concentration of OH^- increases.

In dilute water solutions of acids, the concentration of H^+ is still only small even for strong acids, so it is more convenient to express it via the negative log of the concentration, the pH. Thus instead of referring to the H^+ concentration of a solution as say 0.001 M, it is customary to quote this as its pH which is $-\log_{10}[H^+]$. In this example the pH would be $-\log(0.001) = -\log(10^{-3.0}) = 3.0$. For strong acid solutions more concentrated than about 1 M, pH is not useful. Given that pure water has equal concentrations of H^+ and OH^- and that at $25^\circ C$ this is 1×10^{-7} M, then pure water has a pH = 7.0. Acidic solutions have a pH < 7 and basic solutions have a pH > 7.

The pH of a strong acid solution can be calculated directly from the acid concentration because this is the same as the resultant concentration of H^+ . The pH of a basic solution can be calculated by first calculating the H^+ concentration from the equilibrium expression $[H^+] \times [OH^-] = 1 \times 10^{-14}$ at room temperature.

The pH of weak acids cannot be calculated directly from the concentration of the acid as it is only partially dissociated.

An acid-base indicator is a weak acid which is either colourless or has an intense colour and whose conjugate base has a different colour. As the pH of a solution containing a small amount of indicator is increased, e.g. by adding OH^- solution, the relative amounts of the acid and base form of the indicator change. Ultimately as the weak acid form of the indicator is converted to its conjugate base, the colour of the solution changes to that of the conjugate base. The actual pH at which an indicator changes colour is a characteristic of each indicator and is not necessarily at pH = 7 although for the most commonly known indicator, litmus, the colour change does occur at pH = 7.

TUTORIAL QUESTIONS - TOPIC 13.

1. Explain the meaning of each of the following terms:

conjugate acid/base pair;

diprotic acid;

triprotic acid;

dynamic equilibrium;

strong acid;

weak acid;

self-ionising solvent;

pH.

2. Give the formula for the conjugate base of each of the following acids:

HNO_3

HNO_2

HCN

HF

HClO_4

H_2SO_4

HPO_4^{2-}

HI

CH_3COOH

$\text{CH}_3\text{CH}_2\text{COOH}$

HSO_4^-

H_2S

H_3PO_4

HCOOH

NH_4^+

HS^-

3. Give the formula for the conjugate acid of each of the following bases:

Cl^-

CN^-

SO_4^{2-}

CO_3^{2-}

HCOO^-

I^-

NH_3

O^{2-}

IO_4^-

HSO_3^-

NO_2^-

$\text{C}_2\text{H}_5\text{O}^-$

HS^-

H_2O

OH^-

$\text{C}_6\text{H}_5\text{COO}^-$

4. The following are strong acids in water solution. Write an equation for the ionisation of each in water.

(i) HI

(ii) HBrO_4

(iii) HNO_3

(iv) HCl

5. The following are weak acids in water solution. Write an equation for the equilibrium which is established in each case.

(i) HF

(ii) HCN

(iii) $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}$

(iv) H_2S

6. Calculate the pH of the following solutions of strong acids.

(i) 0.010 M HNO_3

(ii) 0.15 M HI

(iii) 0.05 M HClO_4

(iv) 1.5×10^{-3} M HCl

(v) 5.8×10^{-4} M HBr

(vi) 0.020 M HNO_3

7. Calculate the concentration of H^+ present in solutions having the following pH values.

- (i) 3.10 (ii) 5.21 (iii) 2.66 (iv) 7.00 (v) 1.55 (vi) 4.10

8. (i) Express the following H_3O^+ concentrations as a pH and describe each as either "acidic" or "basic":

- (a) $1.0 \times 10^{-3} M$ (b) $5.4 \times 10^{-9} M$

(ii) Calculate the pH of a $6.3 \times 10^{-6} M$ hydrochloric acid solution.

(iii) Calculate the concentration of H^+ in a solution which has $pH = 8.37$.

9. (a) Calculate the pH of the following sodium hydroxide solutions which contain the strong base, OH^- :

- (i) NaOH (0.010 M) (ii) NaOH ($1.0 \times 10^{-3} M$)

- (iii) NaOH ($1.0 \times 10^{-4} M$) (iv) NaOH ($1.0 \times 10^{-5} M$)

(b) What do you notice about the pH of basic solutions as the concentration of hydroxide ion decreases? How does this compare with the pH of an acid as it is progressively diluted with water?

ANSWERS TO TUTORIAL TOPIC 13

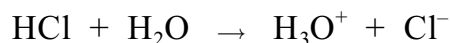
1. Conjugate acid/base pair: Any two species whose formulas differ by a **single** H^+ ion - e.g. HCl and Cl^- where HCl is the conjugate acid of Cl^- and Cl^- is the conjugate base of HCl .

Diprotic acid: An acid that can donate two H^+ ions - e.g. H_2SO_4 .

Triprotic acid: An acid that can donate three H^+ ions - e.g. H_3PO_4 .

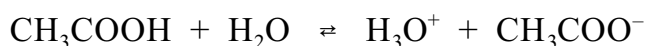
Dynamic equilibrium: The state achieved in a chemical reaction whereby the rate at which the reaction is proceeding from left to right is equal to the rate at which the reverse reaction is proceeding - the net effect is that there is no overall change in the amounts of any of the components involved in the reaction even though they are reacting.

Strong acid: If all of the molecules or ions of a species in solution donate H^+ ions to the solvent, that substance is called a strong acid. The solvent is usually (but not necessarily) water. This is shown by the use of a single arrow. e.g. hydrogen chloride

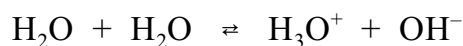


Weak acid: A substance whose molecules or ions do not all react to donate H^+ ions to the solvent. Instead a dynamic equilibrium is established between the acid and its conjugate base. This is shown by use of the equilibrium arrows.

e.g. acetic acid



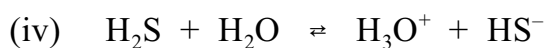
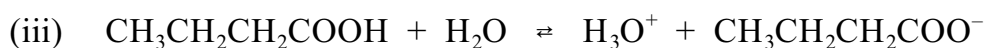
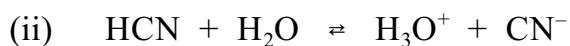
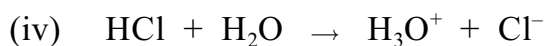
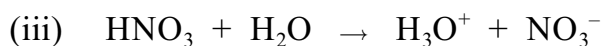
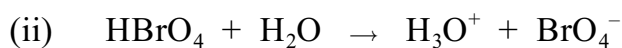
Self ionising solvent: Any species that can both accept and donate an H^+ ion to form the conjugate acid and the conjugate base of that solvent - e.g. water forms H_3O^+ and OH^- when undergoing self ionization.



pH: Defined as $-\log_{10}[\text{H}^+]$ and used to conveniently represent the very small concentrations of H^+ ions present in dilute water solutions of acids and bases. A change of one unit on the pH scale corresponds to a tenfold change in the concentration of H^+ ions.

- | | | | | |
|----|---------------------------|--------------------------------------|--------------------|-----------------|
| 2. | NO_3^- | NO_2^- | CN^- | F^- |
| | ClO_4^- | HSO_4^- | PO_4^{3-} | I^- |
| | CH_3COO^- | $\text{CH}_3\text{CH}_2\text{COO}^-$ | SO_4^{2-} | HS^- |
| | H_2PO_4^- | HCO_2^- | NH_3 | S^{2-} |

3.	HCl	HCN	HSO_4^-	HCO_3^-
	HCOOH	HI	NH_4^+	OH^-
	HIO_4	H_2SO_3	HNO_2	$\text{C}_2\text{H}_5\text{OH}$
	H_2S	H_3O^+	H_2O	$\text{C}_6\text{H}_5\text{COOH}$



6. (i) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.010) = -(-2.00) = 2.00$

(ii) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.15) = -(-0.82) = 0.82$

(iii) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.05) = -(-1.3) = 1.3$

(iv) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(1.5 \times 10^{-3}) = -(-2.82) = 2.82$

(v) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(5.8 \times 10^{-4}) = -(-3.24) = 3.24$

(vi) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.020) = -(-1.70) = 1.70$

7. (i) $\text{pH} = -\log_{10}[\text{H}^+] \quad \therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(3.10)} = 7.9 \times 10^{-4} \text{ M}$

(ii) $\text{pH} = -\log_{10}[\text{H}^+] \quad \therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(5.21)} = 6.2 \times 10^{-6} \text{ M}$

(iii) $\text{pH} = -\log_{10}[\text{H}^+] \quad \therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(2.66)} = 2.2 \times 10^{-3} \text{ M}$

(iv) $\text{pH} = -\log_{10}[\text{H}^+] \quad \therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(7.00)} = 1.0 \times 10^{-7} \text{ M}$

(v) $\text{pH} = -\log_{10}[\text{H}^+] \quad \therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(1.55)} = 2.8 \times 10^{-2} \text{ M}$

(vi) $\text{pH} = -\log_{10}[\text{H}^+] \quad \therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(4.10)} = 7.9 \times 10^{-5} \text{ M}$

8. (i) (a) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(1.0 \times 10^{-3}) = -(-3.00) = 3.00$
 (b) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(5.4 \times 10^{-9}) = -(-8.27) = 8.27$
 (ii) (a) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(6.3 \times 10^{-6}) = -(-5.20) = 5.20$
 (iii) $\text{pH} = -\log_{10}[\text{H}^+]$
 $\therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-(8.37)} = 4.3 \times 10^{-9} \text{ M}$

- 9.(a) (i) As the solutions are those containing the strong base OH^- and it is the concentration of OH^- that is given, the first step is to calculate the concentration of H^+ using the relation $[\text{H}^+] \times [\text{OH}^-] = 10^{-14}$ at 25°C .

$$[\text{H}^+] \times [\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] \times (0.010) = 10^{-14}$$

$$\therefore [\text{H}^+] = 10^{-14} / (0.010) = 10^{-12.00}$$

$$\text{pH} = -\log_{10}(10^{-12.00}) = -(-12.00) = 12.00$$

$$(ii) [\text{H}^+] \times [\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] \times (1.0 \times 10^{-3}) = 10^{-14}$$

$$\therefore [\text{H}^+] = 10^{-14} / (1.0 \times 10^{-3}) = 10^{-11.00}$$

$$\text{pH} = -\log_{10}(10^{-11.00}) = -(-11.00) = 11.00$$

$$(iii) [\text{H}^+] \times [\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] \times (1.0 \times 10^{-4}) = 10^{-14}$$

$$\therefore [\text{H}^+] = 10^{-14} / (1.0 \times 10^{-4}) = 10^{-10.00}$$

$$\text{pH} = -\log_{10}(10^{-10.00}) = -(-10.00) = 10.00$$

$$(iv) [\text{H}^+] \times [\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] \times (1.0 \times 10^{-5}) = 10^{-14}$$

$$\therefore [\text{H}^+] = 10^{-14} / (1.0 \times 10^{-5}) = 10^{-9.00}$$

$$\text{pH} = -\log_{10}(10^{-9.00}) = -(-9.00) = 9.00$$

- (b) When a basic solution is diluted, its pH approaches 7 from above while the pH of an acidic solution approaches 7 from below when it is diluted.