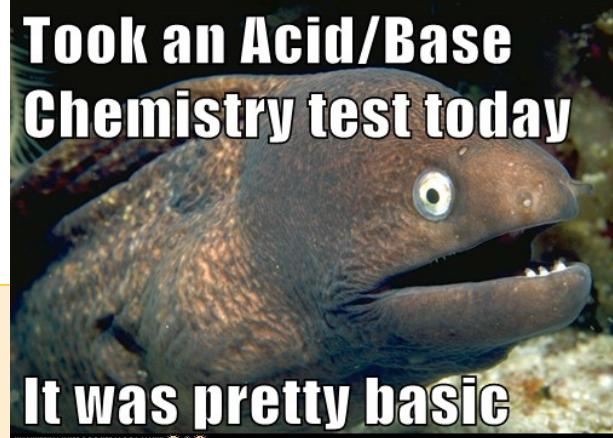


TOPIC 2:

ACIDS & BASES

Syllabus dot-points

Science as a Human Endeavour:



Models of acids and bases:

Models and theories are contested and refined or replaced when new evidence challenges them, or when a new model or theory has greater explanatory scope. Davy initially proposed that acids were substances that contained replaceable hydrogen (hydrogen that could be partly or totally replaced by metals) and bases were substances that reacted with acids to form salts and water. The Arrhenius model, which includes only soluble acids and bases, identified acids as substances which produce hydrogen ions in solution and bases as substances which produce hydroxide ions in solution. Subsequently, the Brønsted-Lowry model describes acid-base behaviour in terms of proton donors and proton acceptors. This approach includes a wider range of substances and can be more broadly applied.

Ocean Acidification:

Levels of carbon dioxide in the atmosphere are rising and have a significant impact on global systems, including surface temperatures. The increasing level of carbon dioxide in the atmosphere causes more carbon dioxide to dissolve in the ocean producing carbonic acid and leading to increased ocean acidity. This is predicted to have a range of negative consequences for marine ecosystems such as coral reefs. Calcification is the process which results in the formation of calcium carbonate structures in marine organisms. Acidification shifts the equilibrium of carbonate chemistry in seawater, decreasing the rate and amount of calcification among a wide range of marine organisms. The United Nations Kyoto Protocol and the Intergovernmental Panel on Climate Change aim to secure a global commitment to reducing greenhouse gas emissions over the next few decades.

Science Understanding:

- acids are substances that can act as proton (hydrogen ion) donors and can be classified as monoprotic or polyprotic, depending on the number of protons available for donation
- the relationship between acids and bases in equilibrium systems can be explained using the Brønsted-Lowry model and represented using chemical equations that illustrate the transfer of protons between conjugate acid-base pairs
- the strength of acids is explained by the degree of ionisation at equilibrium in aqueous solution which can be represented by chemical equations and acidity constants (K_a)
- acid-base indicators are weak acids, or weak bases, in which the acidic form is a different colour from the basic form
- the hydrolysis of salts of weak acids and weak bases can be represented using equations; the Brønsted-Lowry model can be applied to explain the acidic, basic and neutral nature of salts derived from bases and monoprotic and polyprotic acids
- water is a weak electrolyte; the self-ionisation of water is represented by $K_w = [H^+][OH^-]$ where $K_w = 1.0 \times 10^{-14}$ at 25 °C
- K_w can be used to calculate the concentration of hydrogen ions or hydroxide ions in solutions of strong acids or bases
- the pH scale is a logarithmic scale and the pH of a solution can be calculated from the concentration of hydrogen ions using the relationship $pH = -\log_{10} [H^+]$
- buffer solutions are conjugate in nature and resist changes in pH when small amounts of strong acid or base are added to the solution; buffering capacity can be explained qualitatively; Le Châtelier's Principle can be applied to predict how buffers respond to the addition of hydrogen ions and hydroxide ions

Historical Models of Acids & Bases

Boyle (1660s):

- Used litmus to classify acids and bases.
- Defined their characteristic properties.

Acids	Bases (Alkalies)
Sour taste	Feel slippery
Corrosive	
Turns blue litmus red	Turns red litmus blue
Becomes less acidic when combined with bases	Becomes less alkaline when combined with acids

Lavoisier (1770s):

- Studied combustion of carbon, sulfur, phosphorous and nitrogen
- Products of combustion showed acidic properties
- Concluded that acidity is caused by oxygen**

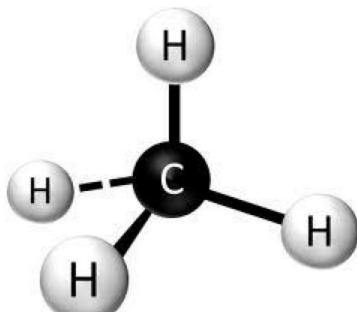
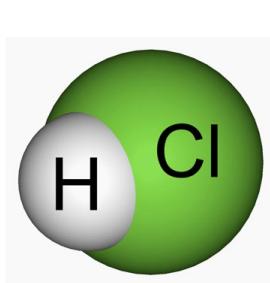
Davy (1810s):

- Discovered the element chlorine
- Showed that the substance 'oxymuriatic acid' (now called hydrochloric acid) didn't contain oxygen
- This contradicted previous model of acids & bases. Davy instead proposed acids contain hydrogen.
- Von Liebig extended Davy's theory by defining acids as substances that contain replaceable hydrogen (hydrogen that can be replaced by metals)



Limitations of Davy's model:

- Does not explain why some substances containing H (e.g. HCl) are acidic but others (CH_4) are not



Arrhenius theory

Main principles:

- Acids produce **H⁺ ions** when dissolved in water
- Bases produce **OH⁻ ions** when dissolved in water
- Acids and bases neutralise each other: $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$

Ionisation / Dissociation equations

- Ionisation refers to a covalent molecular substance forming ions
 - $\text{HC}\ell(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{C}\ell^-(\text{aq})$
 - $\text{CH}_3\text{COOH}(\text{aq}) \rightarrow \text{CH}_3\text{COO}^-(\text{aq}) + \text{H}^+(\text{aq})$
- Dissociation refers to the separation of ions from an ionic substance when dissolved in water
 - $\text{NaOH}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{Ba}(\text{OH})_2(\text{s}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$

Monoprotic acids:

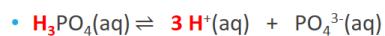
- Only contain one ionisable hydrogen atom

Examples: **HCl**, **HNO₃**, **CH₃COOH**

Polyprotic acids:

- Contain multiple ionisable hydrogen atom
 - Diprotic: **H₂SO₄**, **H₂CO₃**
 - Triprotic: **H₃PO₄**
- Each ionisation step can be shown separately... (*successive ionisation*)
 - $\text{H}_3\text{PO}_4(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{H}_2\text{PO}_4^-(\text{aq})$
 - $\text{H}_2\text{PO}_4^-(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$
 - $\text{HPO}_4^{2-}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$

...or summarised into a single equation



Citric acid ($\text{C}_6\text{H}_8\text{O}_7$) is a weak, triprotic acid. Write successive ionisation equations for citric acid.

.....

.....

.....

Affects how many moles of acid are needed to neutralise a base.

<u>Equation</u>	<u>Moles acid</u>	<u>Moles base needed</u>
$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$	1.00 mol	1.00 mol
$\text{H}_2\text{SO}_4 + 2 \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2 \text{H}_2\text{O}$	1.00 mol	2.00 mol
$\text{H}_3\text{PO}_4 + 3 \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + 3 \text{H}_2\text{O}$	1.00 mol	3.00 mol

If you were provided with 1 mol L⁻¹ of an acid, and had to determine whether it is monoprotic or polyprotic, what could you do experimentally? What other chemicals and equipment would you need?

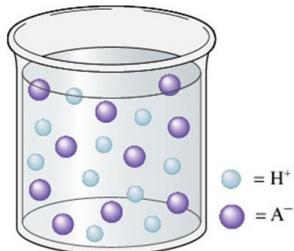
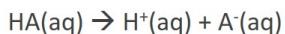
Strong vs Weak

Strong acids fully ionise in solution.

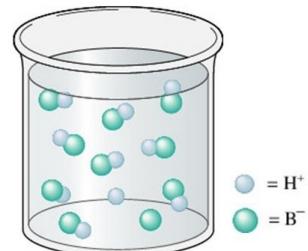
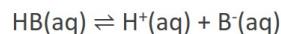
Weak acids only partially ionise in solution. They exist in **equilibrium** with molecules.

Example:

- HA is a strong acid...



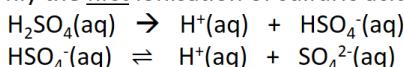
HB is a weak acid...



Common acids:

Strong acids		Weak acids	
Hydrochloric acid	HCl	Ethanoic acid	CH_3COOH
Nitric acid	HNO_3	Carbonic acid	H_2CO_3
Sulfuric acid *	H_2SO_4	Phosphoric acid	H_3PO_4
		Hydrofluoric acid	HF
		Sulfurous acid	H_2SO_3

* Note: Only the first ionisation of sulfuric acid is strong. The second ionisation is weak.



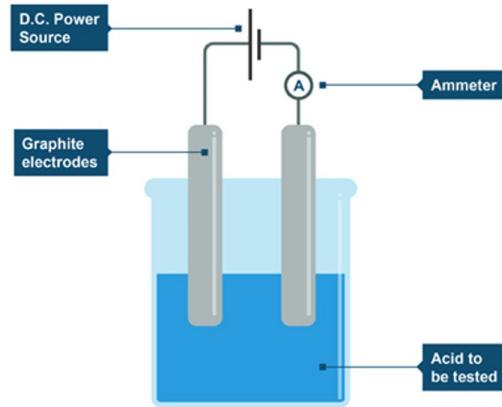
Common bases:

Strong bases		Weak bases	
Metal hydroxides. e.g.	NaOH	Ammonia	NH_3
	KOH	Methylamine	CH_3NH_2
	Ba(OH)_2	Carbonate ion	CO_3^{2-}

Evidence of Arrhenius Theory

Conductivity:

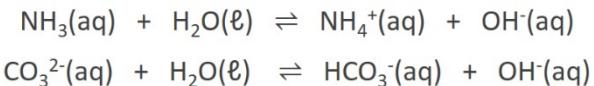
- Solutions can be classified as:
 - Strong electrolytes (conduct electricity well)
 - Weak electrolytes (conduct electricity poorly)
 - Non-electrolytes (do not conduct electricity)
- Conduction of electricity is related to presence of mobile charged particles (e.g. individual ions)
- All solutions of acids and bases conduct electricity,
∴ all solutions of acids and bases contain free ions



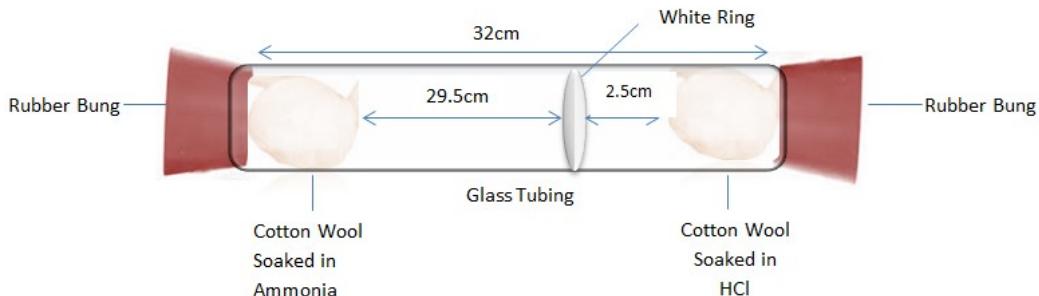
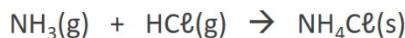
Problems with Arrhenius theory

- * Some substances such as ammonia (NH_3) and sodium carbonate (Na_2CO_3) were known to be basic, but didn't contain OH in their formula.

Some scientists argued that Arrhenius theory was still valid, because these substances could react with water to form OH^- ions.



However, pure gaseous ammonia reacts with HCl even when there is no water present.



- * According to Arrhenius theory, acids and bases should produce neutral solutions, because the H^+ ions will react with the OH^- ions to produce water.
 - e.g. $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\ell)$ $\text{NaCl}(\text{aq})$ is neutral ✓

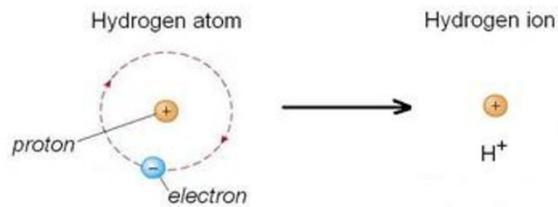
However some acid-base reactions produce **basic solutions**

- e.g. $\text{CH}_3\text{COOH}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCH}_3\text{COO}(\text{aq}) + \text{H}_2\text{O}(\ell)$ $\text{NaCH}_3\text{COO}(\text{aq})$ is basic, not neutral

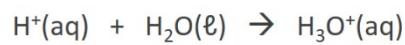
And some acid-base reactions produce **acidic solutions**

- e.g. $\text{HCl}(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4\text{Cl}(\text{aq})$ $\text{NH}_4\text{Cl}(\text{aq})$ is acidic, not neutral

* H^+ represents an individual proton.



In reality, $\text{H}^+(\text{aq})$ would be too unstable to exist in solution for significant amounts of time. It would quickly react with water to form **hydronium ion (H_3O^+)**.



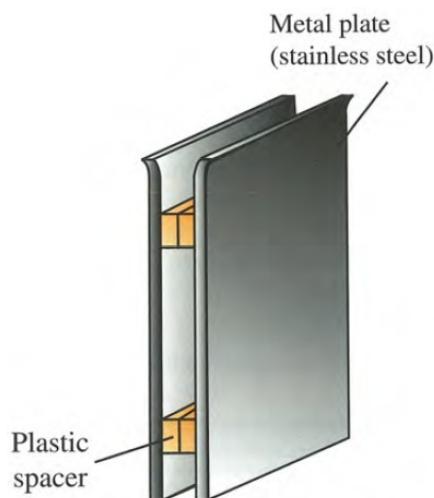
EXPERIMENT: CONDUCTIVITY OF ACIDS AND BASES

Background: Acids and bases are electrolytes, that is they conduct an electric current when in aqueous solution. For acids and bases electrolyte strength is also referred to as acid or base strength. Strong acids and bases exist essentially as ions in aqueous solution. Weak acids and bases are those in which only a small proportion of the molecules or ions react with water to form hydronium ions, H_3O^+ (hydrogen ion attached to a water molecule) or hydroxide ions, OH^- in aqueous solution.

Aim: To investigate the electrical conductivity of acid and base solutions of different strengths and concentrations.

Equipment:

- Power supply
- Plate electrode system
- Ammeter
- Electrical leads (three)
- Two beakers
- 50 mL samples of the following solutions:
 - Hydrochloric acid 1 mol L^{-1}
 - Hydrochloric acid 0.1 mol L^{-1}
 - Hydrochloric acid 0.01 mol L^{-1}
 - Hydrochloric acid 0.001 mol L^{-1}
 - Nitric acid 1 mol L^{-1}
 - Acetic acid 1 mol L^{-1}
 - Sodium hydroxide 1 mol L^{-1}
 - Ammonia 1 mol L^{-1}



Method:

1. Set up the circuit and the electrodes as shown to the right.
Set the power supply to **6 V DC**.
2. Place 50 mL of 1 mol L^{-1} HCl in a beaker, put the electrodes in the solution and briefly turn on the power pack.
3. Record the reading on the ammeter.
4. Wash out the beaker and electrodes with distilled water and then in turn test the conductivity of the other solutions.



Figure 21.2

Results:

Acid	Concentration (mol L^{-1})	Acid/base strength (strong or weak?)	Ammeter reading
Hydrochloric acid (HCl)	1		
	0.1		
	0.01		
	0.001		
Nitric acid (HNO_3)	1		
Acetic acid (CH_3COOH)	1		
Sodium hydroxide (NaOH)	1		
Ammonia (NH_3)	1		

Questions:

1. Explain why these solutions conduct an electric current

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2. Look at your results for the HCl solutions. What conclusion can you draw regarding the conductivity of a solution and the concentration of ions in it?

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3. Were the conductivities of 1.0 mol L^{-1} HCl and 1.0 mol L^{-1} CH_3COOH different? If so, explain why they were different.

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4. Were the conductivities of 1.0 mol L^{-1} NaOH and 1.0 mol L^{-1} NH_3 different? If so, explain why they were different.

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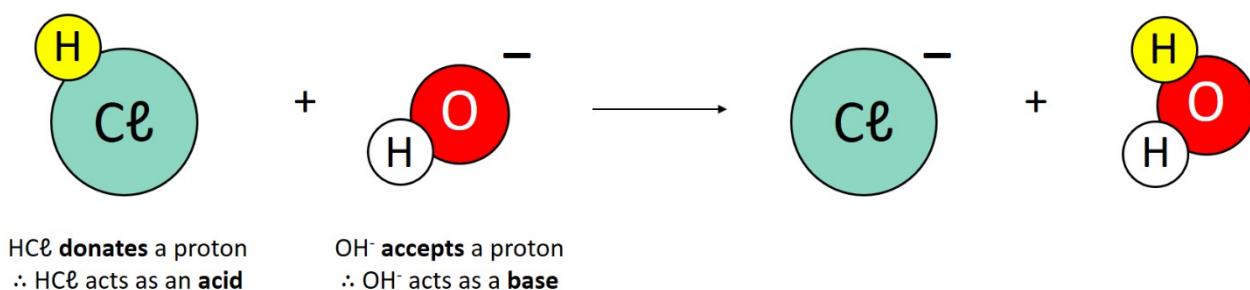
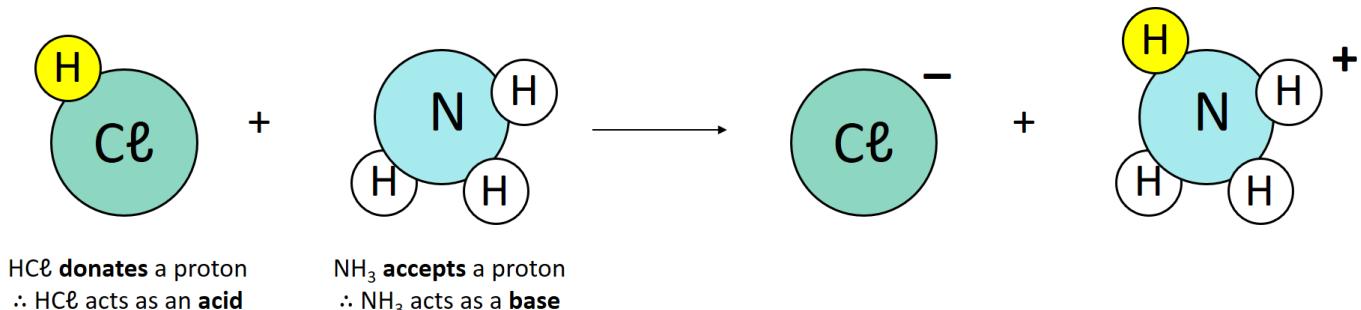
5. Acids and bases are referred to as being 'strong' or 'weak' as distinct from being 'concentrated' or 'dilute'. Explain how these two variables (electrolyte strength and solution concentration) affect the conductivity of aqueous solutions.

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Brønsted-Lowry Theory

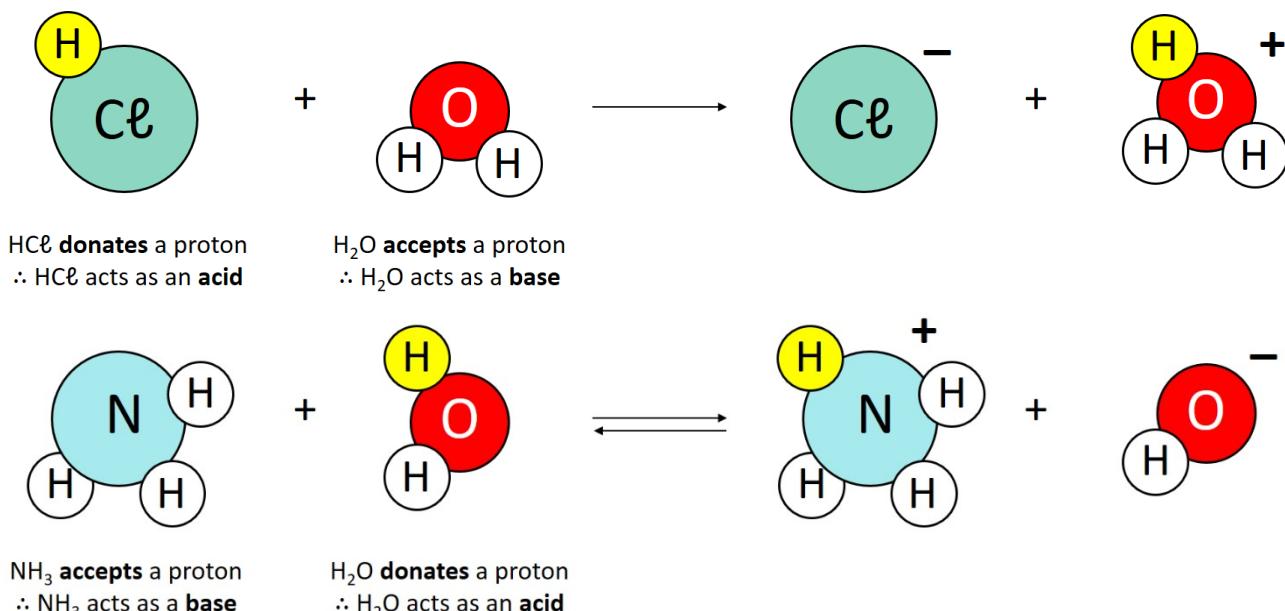
Definitions:

- Acids are substances that **donate** one or more **protons** ('**proton donors**')
- Bases are substances that **accept** one or more **protons** ('**proton acceptors**')



Acids and bases in aqueous solutions:

Brønsted-Lowry theory also applies to acids and bases dissolving in water to form an aqueous solution. Water will either accept a proton to form H₃O⁺(aq) or donate a proton to form OH⁻.



Substances which can either accept or donate protons such as water are called **amphiprotic** or **amphoteric**.

IMPORTANT! This affects how we write ionisation equations for acids.

Arrhenius theory:



Brønsted-Lowry theory:



Question:

- Write an ionisation equation using Arrhenius theory
- Write an ionisation equation using Brønsted-Lowry theory

	Arrhenius theory:
Nitric acid	Brønsted-Lowry theory:

	Arrhenius theory:
Hydrofluoric acid	Brønsted-Lowry theory:

Carbonate ion (weak base)	Brønsted-Lowry theory:
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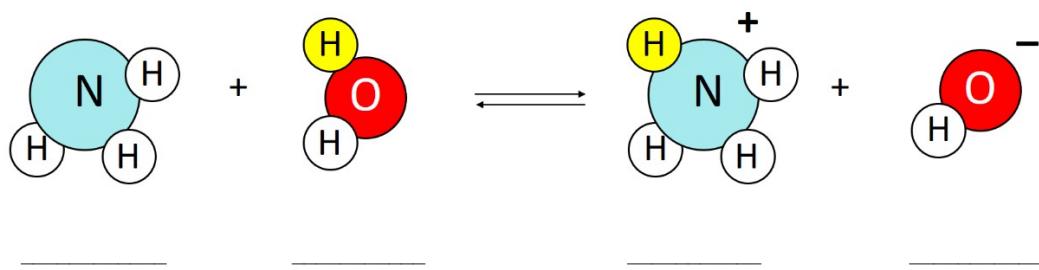
Methylamine (CH_3NH_2) (weak base)	Brønsted-Lowry theory:
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Identify the Brønsted-Lowry acid and bases. Show reasoning.

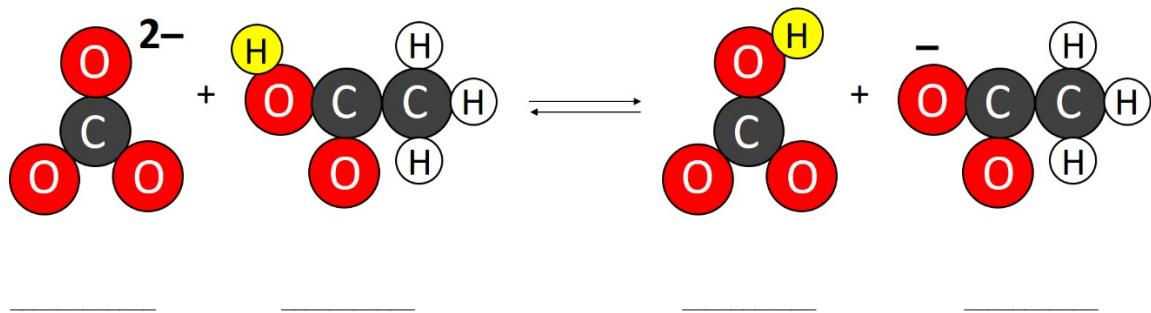


Conjugate acids and bases

Recall that ammonia is a weak base, hence its ionisation is a **reversible reaction**.



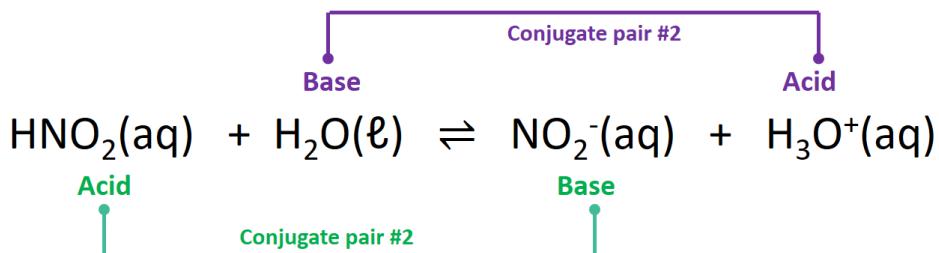
Reaction between acetic acid and sodium carbonate:



In a reversible reaction:

- The species produced from a **weak acid** will be **basic**
- The species produced from a **weak base** will be **acidic**

These groups are referred to as **conjugate acid-base pairs**



For the above equation it can also be said that:

- NO_2^- is the conjugate base of HNO_2
- HNO_2 is the conjugate acid of NO_2^-

H_2O is the conjugate base of H_3O^+
 H_3O^+ is the conjugate acid of H_2O

An acid and its conjugate base will only differ in formula by one H^+ ion.

Brønsted-Lowry review questions:

1. Which of the following reactions are regarded as Brønsted-Lowry acid-base reactions?

For these reactions identify which reactant is acting as the Brønsted-Lowry acid and which is acting as the Brønsted-Lowry base.

- $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HSO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
- $\text{HS}^-(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{S}^{2-}(\text{aq})$
- $\text{H}_2\text{O}_2(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\ell) + \text{O}_2(\text{g})$
- $\text{H}_2\text{O}(\ell) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$

2. Write equations for the dihydrogenphosphate ion reacting with water as a:

- Brønsted-Lowry acid

.....

- Brønsted-Lowry base

.....

3. The hydrogensulfide ion, HS^- , is amphiprotic.

- What is meant by the word 'amphiprotic'?

.....

- When HS^- is added to hydrochloric acid, it acts as a Brønsted-Lowry base. Write the equation for the reaction between $\text{HS}^-(\text{aq})$ and the hydronium ions in the hydrochloric acid.

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- HS^- acts as a Brønsted-Lowry acid when it is added to a solution of sodium hydroxide. Write an equation for the reaction between $\text{HS}^-(\text{aq})$ and the hydroxide ions in the sodium hydroxide solution

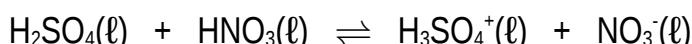
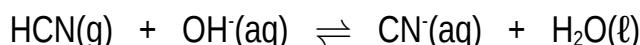
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4. Complete the tables by giving the formula of the conjugate acids/bases.

Acid	Conjugate base
HI	
HNO_2	
HCO_3^-	
PH_4^+	

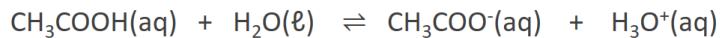
Base	Conjugate acid
SO_3^{2-}	
O^{2-}	
HF	
NH_3	

5. Draw lines linking the conjugate acid-base pairs in the following reactions. Label each species as being an acid or a base.



Chemical equilibrium

Consider a solution of ethanoic acid in equilibrium:

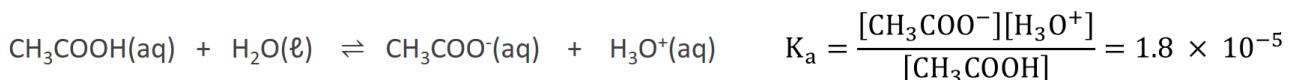


This solution would be in chemical equilibrium, therefore we can apply principles from the equilibrium topic.

- While in equilibrium, the rate of forward and reverse reactions would be equal
- While in equilibrium, the concentration of all species would remain constant
- Changing the concentration of a species would affect equilibrium.
e.g. according to Le Châtelier's Principle, removing H_3O^+ would cause reaction to shift to the right

As with other systems, we can write **equilibrium constant expressions** to show the relative concentrations of species at equilibrium.

For the ionisation of an acid, this value is called the **acid dissociation constant (K_a)**.



1. What does the size of the acid dissociation constant (K_a) for CH_3COOH indicate about acetic acid?

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2. What do the K_a values of CH_3COOH and HF indicate about their *relative* strengths? Justify your answer.

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RELATIVE STRENGTHS OF BRØNSTED-LOWRY ACIDS AND BASES

in aqueous solution at room temperature.

Name of Acid

Acid

Base

K_a

STRONG ↑

Perchloric	$\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$		very large
Hydriodic	$\text{HI} \rightarrow \text{H}^+ + \text{I}^-$		very large
Hydrobromic	$\text{HBr} \rightarrow \text{H}^+ + \text{Br}^-$		very large
Hydrochloric	$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$		very large
Nitric	$\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$		very large
Sulphuric	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$		very large
Hydronium Ion	$\text{H}_3\text{O}^+ \rightleftharpoons \text{H}^+ + \text{H}_2\text{O}$		1.0
Iodic	$\text{HIO}_3 \rightleftharpoons \text{H}^+ + \text{IO}_3^-$		1.7×10^{-1}
Oxalic	$\text{H}_2\text{C}_2\text{O}_4 \rightleftharpoons \text{H}^+ + \text{HC}_2\text{O}_4^-$		5.9×10^{-2}
Sulphurous ($\text{SO}_2 + \text{H}_2\text{O}$)	$\text{H}_2\text{SO}_3 \rightleftharpoons \text{H}^+ + \text{HSO}_3^-$		1.5×10^{-2}
Hydrogen sulphate ion	$\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2-}$		1.2×10^{-2}
Phosphoric	$\text{H}_3\text{PO}_4 \rightleftharpoons \text{H}^+ + \text{H}_2\text{PO}_4^-$		7.5×10^{-3}
Hexaaquoiron ion, iron(III) ion	$\text{Fe}(\text{H}_2\text{O})_6^{3+} \rightleftharpoons \text{H}^+ + \text{Fe}(\text{H}_2\text{O})_5(\text{OH})^{2+}$		6.0×10^{-3}
Citric	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7 \rightleftharpoons \text{H}^+ + \text{H}_2\text{C}_6\text{H}_5\text{O}_7^-$		7.1×10^{-4}
Nitrous	$\text{HNO}_2 \rightleftharpoons \text{H}^+ + \text{NO}_2^-$		4.6×10^{-4}
Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$		3.5×10^{-4}
Methanoic, formic	$\text{HCOOH} \rightleftharpoons \text{H}^+ + \text{HCOO}^-$		1.8×10^{-4}
Hexaaquochromium ion, chromium(III) ion	$\text{Cr}(\text{H}_2\text{O})_6^{3+} \rightleftharpoons \text{H}^+ + \text{Cr}(\text{H}_2\text{O})_5(\text{OH})^{2+}$		1.5×10^{-4}
Benzoic	$\text{C}_6\text{H}_5\text{COOH} \rightleftharpoons \text{H}^+ + \text{C}_6\text{H}_5\text{COO}^-$		6.5×10^{-5}
Hydrogen oxalate ion	$\text{HC}_2\text{O}_4^- \rightleftharpoons \text{H}^+ + \text{C}_2\text{O}_4^{2-}$		6.4×10^{-5}
Ethanoic, acetic	$\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$		1.8×10^{-5}
Dihydrogen citrate ion	$\text{H}_2\text{C}_6\text{H}_5\text{O}_7^- \rightleftharpoons \text{H}^+ + \text{HC}_6\text{H}_5\text{O}_7^{2-}$		1.7×10^{-5}
Hexaaquoaluminum ion, aluminum ion	$\text{Al}(\text{H}_2\text{O})_6^{3+} \rightleftharpoons \text{H}^+ + \text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+}$		1.4×10^{-5}
Carbonic ($\text{CO}_2 + \text{H}_2\text{O}$)	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$		4.3×10^{-7}
Monohydrogen citrate ion	$\text{HC}_6\text{H}_5\text{O}_7^{2-} \rightleftharpoons \text{H}^+ + \text{C}_6\text{H}_5\text{O}_7^{3-}$		4.1×10^{-7}
Hydrogen sulphite ion	$\text{HSO}_3^- \rightleftharpoons \text{H}^+ + \text{SO}_3^{2-}$		1.0×10^{-7}
Hydrogen sulphide	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$		9.1×10^{-8}
Dihydrogen phosphate ion	$\text{H}_2\text{PO}_4^- \rightleftharpoons \text{H}^+ + \text{HPO}_4^{2-}$		6.2×10^{-8}
Boric	$\text{H}_3\text{BO}_3 \rightleftharpoons \text{H}^+ + \text{H}_2\text{BO}_3^-$		7.3×10^{-10}
Ammonium ion	$\text{NH}_4^+ \rightleftharpoons \text{H}^+ + \text{NH}_3$		5.6×10^{-10}
Hydrocyanic	$\text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^-$		4.9×10^{-10}
Phenol	$\text{C}_6\text{H}_5\text{OH} \rightleftharpoons \text{H}^+ + \text{C}_6\text{H}_5\text{O}^-$		1.3×10^{-10}
Hydrogen carbonate ion	$\text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-}$		5.6×10^{-11}
Hydrogen peroxide	$\text{H}_2\text{O}_2 \rightleftharpoons \text{H}^+ + \text{HO}_2^-$		2.4×10^{-12}
Monohydrogen phosphate ion	$\text{HPO}_4^{2-} \rightleftharpoons \text{H}^+ + \text{PO}_4^{3-}$		2.2×10^{-13}
Water	$\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$		1.0×10^{-14}
Hydroxide ion	$\text{OH}^- \leftarrow \text{H}^+ + \text{O}^{2-}$		very small
Ammonia	$\text{NH}_3 \leftarrow \text{H}^+ + \text{NH}_2^-$		very small

WEAK

STRENGTH OF BASE

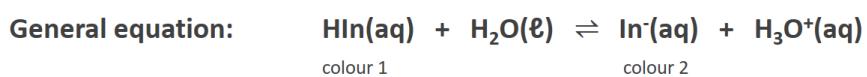
STRONG ↓

Indicators

Indicators are chemical species that change colour in different pH conditions.

Most indicators are an acid-base conjugate pair, where each species in the pair is a different colour.

(Use a highlighter to show appropriate colour for each indicator)

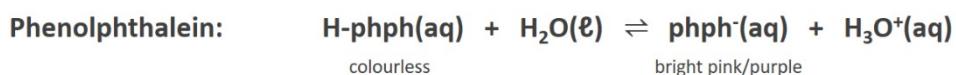


In the above equation, “In” is used as shorthand for “indicator”. The indicator molecule can either be protonated (HIn) or de-protonated (In^-).

Assume that there is initially an equal mix of HIn(aq) and $\text{In}^-(\text{aq})$.

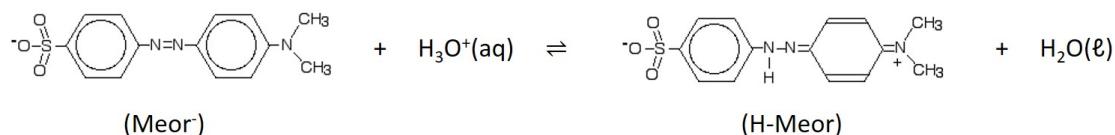
According to Le Châtelier's principle, what would happen to the colour if...

- HCl(aq) was added?
 - NaOH(aq) was added?



Explain using Le Châtelier's principle why phenolphthalein turns purple in solutions with pH > 9.

Methyl orange indicator:

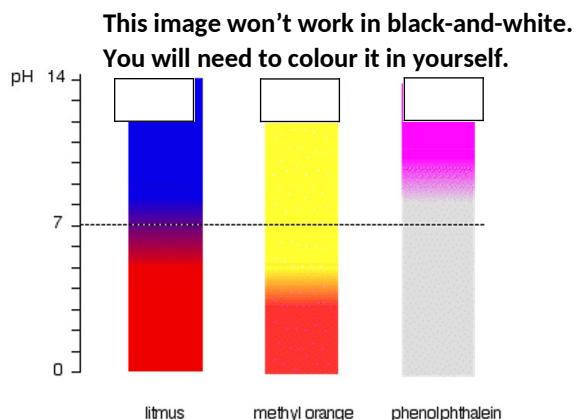


Using the methyl orange photograph on the previous page, determine the colours of MeO^- and $\text{H-}\text{MeO}$. Justify your decision using Le Châtelier's principle.

Most indicators only have 2 or 3 different coloured forms.

Indicators do not necessarily change at pH 7. Their changing point depends on the equilibrium constant for that indicator.

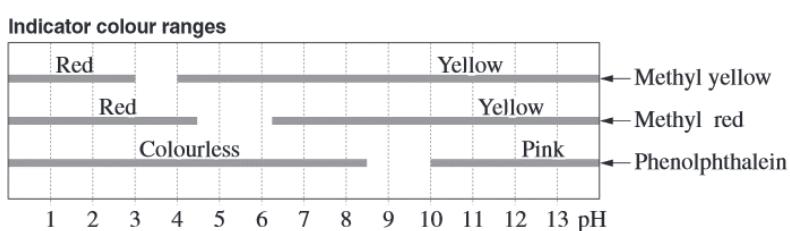
Universal indicator is made of a mixture of four different types of indicator. This allows it to display a wide range of colours across the pH scale.



Q: To determine the pH of garden soil, a sample was first saturated with distilled water in a petri dish. Drops of three indicators listed below were added to separate parts of the sample. The colours observed are shown in the table.

Experimental results

<i>Indicator</i>	Methyl yellow	Methyl red	Phenolphthalein
<i>Colour observed</i>	Yellow	Red	Colourless



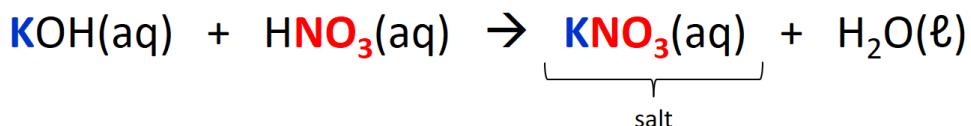
Plant response	
Plant	soil pH range for optimal growth
Carrot	5.5 – 6.8
Chrysanthemum	6.0 – 6.3
Hydrangea Blue	4.0 – 5.0
Hydrangea White	6.5 – 8.0
Potato	5.0 – 5.7

Using the information given, select the plant that will grow well at the current soil pH, and justify your selection.

Salts from acid-base reactions

Salts are ionic compounds produced after the reaction of an acid and a base.

In general, the cation of the salt comes from the base and the anion comes from the acid.



Not all salts are neutral. Salts may be acidic or basic!

Acidic salts

Acidic salts contain an ion which is a **proton donor**.

Example:

Ammonium chloride is an acidic salt because ammonium ion (NH_4^+) is a proton donor.

When dissolved in water, the NH_4^+ ion undergoes a hydrolysis reaction.
It donates a proton to water molecules to form H_3O^+ .



This increase the $[\text{H}_3\text{O}^+]$ relative to $[\text{OH}^-]$, therefore decreases the pH of the solution.

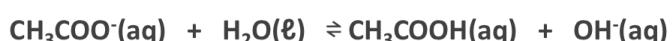
Basic salts

Basic salts contain an ion which is a **proton acceptor**.

Example:

Sodium acetate is a basic salt because acetate ion (CH_3COO^-) is a proton acceptor.

When dissolved in water, the CH_3COO^- ion undergoes a hydrolysis reaction.
It accepts a proton to water molecules to form OH^- .



This increase the $[\text{OH}^-]$ relative to $[\text{H}_3\text{O}^+]$, therefore increases the pH of the solution.

Neutral salts

Neutral salts do not contain protons acceptors or donors.

Example:

Sodium chloride is a neutral salt.

Neither ion undergoes hydrolysis with water to accept or donate protons.

****BE CAREFUL****

Some students **incorrectly** assume Cl^- could accept protons to form HCl

This isn't true because HCl is a **strong acid**. $\text{HCl} + \text{H}_2\text{O} \rightarrow \text{Cl}^- + \text{H}_3\text{O}^+$

If Cl^- *could* accept protons to form HCl then the above would be a reversible reaction and HCl would be a **weak acid**

Shortcut...

You may have noticed a pattern...

Basic salts (e.g. NaCH_3COO) contain an ion from a *weak acid*

Acidic salts (e.g. NH_4Cl) contain an ion from a *weak base*

Neutral salts (e.g. NaOH) contain ions from strong acids and bases

This can be used as a shortcut to quickly figure out if a salt is acidic, basic or neutral.

But it is not an **explanation**.

If it asks '*why is sodium acetate a basic salt*' you can't say '*because acetic acid is a weak acid*'. You would have to talk about the acetate ion accepting a proton, give the equation and state how it affects the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ in solution.

Classify the following salts as acidic, basic or neutral. If it is acidic or basic, include an appropriate hydrolysis equation.

Salt	Acidic, basic or neutral	Hydrolysis equation
NaNO_3		
NaF		
NH_4NO_3		
KI		
Na_3PO_4		
CaCO_3		
MgBr		
NaClO	Basic	

Salts from polyprotic acids

If a salt is produced from a **polyprotic acid**, then the resulting salt might be acidic **or** basic, depending on which acid is involved.

Consider the salt sodium hydrogencarbonate (NaHCO_3). When dissolved in water, it will first dissociate and then the HCO_3^- can react with water. But it can do so in two different ways...



or



The final pH of the solution depends on which of the above reactions occurs to the greatest extent. This could be compared using **equilibrium constants**.

$$K_a = \frac{[\text{CO}_3^{2-}][\text{H}_3\text{O}^+]}{[\text{HCO}_3^-]} = 4.8 \times 10^{-11}$$

$$K_b = \frac{[\text{H}_2\text{CO}_3][\text{OH}^-]}{[\text{HCO}_3^-]} = 2.3 \times 10^{-8}$$

Based on the K_a and K_b values, would the HCO_3^- form an acidic or basic solution? Explain.

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Unless provided with K_a and K_b values in a test, there is no way to determine whether an ion like HCO_3^- would form an acidic or basic solution.

That means that you will have to **memorise** the following table so you can correctly remember which salts from polyprotic acids are acidic, basic and neutral.

Related acid	Acidic ions (will donate proton and form H_3O^+)	Basic ions (will accept proton and form OH^-)
Sulfuric acid	HSO_4^-	SO_4^{2-}
Carbonic acid		HCO_3^- CO_3^{2-}
Phosphoric acid	H_2PO_4^-	HPO_4^{2-} PO_4^{3-}

Salts from the first ionisation of sulfuric and phosphoric are acidic

If you can't remember in a test, guess basic

Questions:

1. a) Predict whether a sample of sodium hydrogensulfate (NaHSO_4) will be acidic or basic.

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- b) Write a hydrolysis equation showing the reaction that results in the above acidity or basicity.

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- c) Use the equation above to explain why the pH of water changes when NaHSO_4 is dissolved in it.

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2. Potassium hydrogenphosphate (Na_2HPO_4) is basic, whereas potassium dihydrogenphosphate (NaH_2PO_4) is acidic. Use the concept of acid and base dissociation constants to explain this difference.

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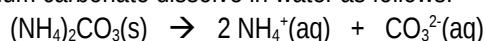
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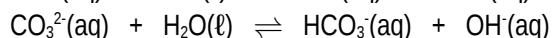
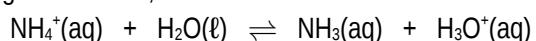
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3. Ammonium carbonate dissolve in water as follows:



Following dissociation, two further reactions occur as follows:



Indicators show that an aqueous solution of $(\text{NH}_4)_2\text{CO}_3$ is basic. Explain this fact in view of the fact that there are twice as many NH_4^+ ions (which may produce H_3O^+) as there are CO_3^{2-} ions (which may produce OH^-) in any sample of ammonium carbonate.

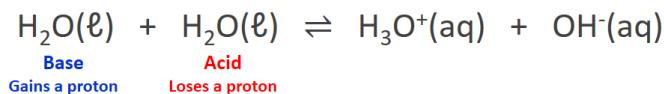
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Water as an electrolyte

Water is a weak electrolyte. It can self-ionise to produce OH^- and H^+ .



As with other equilibrium reactions we can write an equation for the equilibrium constant.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

At 25 °C, $K_w = 1.0 \times 10^{-14}$. This is a very small number and indicates that the reaction heavily favours the reactants. The concentration of H^+ and OH^- in neutral water is very low.

Calculating $[\text{H}^+]$ and $[\text{OH}^-]$

Note: In these calculations, H^+ and H_3O^+ can be used interchangeably. H^+ (from Arrhenius theory) is usually shown in calculations because it is faster and easier than writing H_3O^+ .

If we know $[\text{H}^+]$ in a solution, then we can calculate $[\text{OH}^-]$ using the value of K_w .

$$K_w = [\text{H}^+][\text{OH}^-]$$

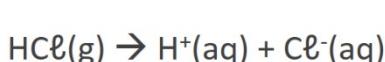
$$\therefore [\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1 \times 10^{-14}}{[\text{H}^+]}$$

e.g. A solution of acid contains 1.5×10^{-5} mol L⁻¹ of H^+ . Calculate $[\text{OH}^-]$.

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1 \times 10^{-14}}{1.5 \times 10^{-5}} = 6.67 \times 10^{-10} \text{ mol/L}$$

Example:

Calculate the concentration of H^+ and OH^- in a 1.0×10^{-2} mol L⁻¹ HCl solution at 25 °C.



$$K_w = [\text{H}^+][\text{OH}^-]$$

$$[\text{H}^+] = [\text{HCl}]$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]}$$

$$= 1.0 \times 10^{-2} \text{ mol L}^{-1}$$

$$= \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}}$$

$$= 1.0 \times 10^{-12} \text{ mol L}^{-1}$$

Note that we can also rearrange the K_w equation to find $[\text{H}^+]$ if we know $[\text{OH}^-]$.

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$\therefore [\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{[\text{OH}^-]}$$

Questions – Calculating $[H^+]$ and $[OH^-]$

1. The concentration of acid and base in a swimming pool is critical to maintaining an adequate level of chlorine in the water. The concentration of hydrogen ions in the water of an olympic swimming pool was measured to be $1.48 \times 10^{-7} \text{ mol L}^{-1}$. Calculate the concentration of hydroxide ions.

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2. Water in the tailings dam of a gold processing plant was found to have a hydroxide ion concentration of $1.58 \times 10^{-4} \text{ mol L}^{-1}$. Calculate $[H^+]$.

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3. Calculate the hydrogen ion concentration in the following solutions.

- (a) A standard $1.55 \times 10^{-4} \text{ mol L}^{-1}$ sodium hydroxide solution used for analysing the acid content of wine.

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- (b) A $3.90 \times 10^{-2} \text{ mol L}^{-1}$ barium hydroxide solution used to make soap. **Careful! This calculation involves a sneaky trick!**

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4. A 0.125 mol L^{-1} sulfuric acid solution is used to dissolve samples of malachite. Assume that the first hydrogen is fully ionised and the second hydrogen is 10.0% ionised. Calculate the concentration of hydroxide ions in the solution.

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pH scale

pH is calculated using the following formula:

$$\mathbf{pH = -\log[H^+]}$$

...where $[H^+]$ is the concentration of H^+ in mol L⁻¹.

The higher the concentration of H^+ , the lower the value for pH.

You can also do the inverse of this equation to find $[H^+]$ of a solution with known pH.

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

Calculating pH – Strong acids

To calculate the pH of a strong acid solution...

1. Find $[H^+]$
2. Perform the pH calculation

e.g. Calculate the pH of a 0.06 mol L⁻¹ solution of HCl.

$$[H^+] = [HCl] = 0.06 \text{ mol L}^{-1}$$

$$\begin{aligned}\mathbf{pH} &= -\log[H^+] \\ &= -\log[0.06] \\ &= 1.22\end{aligned}$$

Q: Calculate the pH of a 5.0×10^{-3} mol L⁻¹ solution of HNO₃

Q: 2 litres of hydrogen chloride gas (HCl) is collected at **S.T.P.** and is dissolved in 500 L of water. Calculate the pH of the resulting solution.

Calculating pH – Strong bases

There are two ways to go about this. Both methods rely on the fact that $[OH^-]$ and $[H^+]$ are linked. Remember, $K_w = [H^+][OH^-]$

Method 1 (longer):

1. Find $[OH^-]$
2. Find $[H^+]$ using the relationship with K_w $[H^+] = K_w / [OH^-]$
3. Find pH using the formula $pH = -\log[H^+]$

Example (Method 1): Calculate the pH of a $2.0 \times 10^{-2} \text{ mol L}^{-1}$ NaOH solution.



1

$$[OH^-] = [NaOH] \\ = 2.0 \times 10^{-2} \text{ mol L}^{-1}$$

$$2 \quad [H^+] = \frac{K_w}{[OH^-]}$$

$$= \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-2}} \\ = 5.0 \times 10^{-13} \text{ mol L}^{-1}$$

3

$$pH = -\log[H^+] = -\log[5.0 \times 10^{-13}] = 12.3$$

Q: Calculate the pH of the following solution using **Method 1** shown above:

- a) A $0.0065 \text{ mol L}^{-1}$ solution of KOH

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- b) A $6.5 \times 10^{-4} \text{ mol L}^{-1}$ solution of Ba(OH)_2

Be careful when calculating $[OH^-]$

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Calculating pH – Strong bases

Just as pH is a measure of the concentration of H^+ , pOH is a measure of the concentration of OH^-

$$pOH = -\log[OH^-]$$

Due to the equilibrium constant of water, there is a relationship between pH and pOH.

$$pH + pOH = 14$$

This allows us to calculate pH of basic solutions using a shorter method.

Method 2 (shorter):

1. Find $[OH^-]$
2. Find pOH using the formula $pOH = -\log[OH^-]$
3. Find pH using the formula $pH = 14 - pOH$

Example (Method 2): Calculate the pH of a 2.0×10^{-2} mol L⁻¹ NaOH solution.



1

$$\begin{aligned}[OH^-] &= [NaOH] \\ &= 2.0 \times 10^{-2} \text{ mol L}^{-1}\end{aligned}$$

2

$$pOH = -\log[OH^-] = -\log[2.0 \times 10^{-2}] = 1.70$$

3

$$pH = 14 - pOH = 14 - 1.70 = 12.3$$

Q: Calculate the pH of the following solution using **Method 2** shown above:

- a) A 0.000024 mol L⁻¹ solution of KOH

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- b) A 7.5×10^{-2} mol L⁻¹ solution of Ba(OH)₂ Be careful when calculating [OH]⁻

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Questions – pH calculations

1. Calculate the $[H^+]$, $[OH^-]$ and pH of the following solutions.

a. A $0.00500 \text{ mol L}^{-1}$ HNO_3 solution used to reduce the pH of nutrient solutions used for hydroponics

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b. A $0.0100 \text{ mol L}^{-1}$ standard $NaOH$ solution used to analyse vinegar solutions

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c. A 2.00 mol L^{-1} HCl solution used to dissolve limestone prior to analysis

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2. Calculate the $[H^+]$ and $[OH^-]$ of solutions with the following pH:

a. Lemon juice of pH 3.00

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b. Dish washing solution of pH 11.0

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Calculating pH – Mixing two acids

Some problems might involve two acidic solutions with different pH values being mixed together.

To solve:

1. Find $n(H^+)$ in the *first solution* $n(H^+)_1$
 2. Find $n(H^+)$ in the *second solution* $n(H^+)_2$
 3. Find the total $n(H^+)$ ions $n(H^+)_{\text{total}} = n(H^+)_1 + n(H^+)_2$
 4. Find the concentration of the total $n(H^+)$. Remember to use the volume of the **total solution**
 5. Calculate pH as per previous pages

Example: 100 mL of 0.005 mol L⁻¹ HCl is added to 200 mL of 0.0001 mol L⁻¹ HCl

$$\begin{aligned} \textcircled{1} \quad n(H^+)_1 &= c \times V \\ &= 0.005 \times 0.100 \\ &= 0.0005 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(H^+)_2 &= c \times V \\ &= 0.0001 \times 0.200 \\ &= 0.0002 \text{ mol} \end{aligned}$$

$$\begin{aligned} \textcircled{3} \quad n(H^+)_{\text{total}} &= n(H^+)_1 + n(H^+)_2 \\ &= 0.0005 + 0.0002 \\ &= 0.0007 \text{ mol} \end{aligned}$$

4 Total volume = 300 mL = 0.3 L

$$\begin{aligned} c(H^+)_{\text{total}} &= n / V \\ &= 0.0007 / 0.300 \\ &= 0.002333 \text{ mol L}^{-1} \end{aligned}$$

5 $\text{pH} = -\log[\text{H}^+]$
= $-\log[0.002333]$
= 2.63

Question – Mixing acids

A brick cleaner needs to replenish the hydrochloric acid solution he is using. Using a pH meter he measures the pH of his depleted solution to be 2.00. He decides to add 3.00 L of 3.00 mol L⁻¹ HCl solution to 2.00 L of his solution. Calculate the pH of the new solution.

Calculating pH – Mixing two bases

Perform this calculation in the same way as mixing two acids, only now you will be finding $n(OH^-)$.

To solve:

1. Find $n(OH^-)$ in the *first solution* $n(OH^-)_1$
2. Find $n(OH^-)$ in the *second solution* $n(OH^-)_2$
3. Find the total $n(OH^-)$ ions $n(OH^-)_{total} = n(OH^-)_1 + n(OH^-)_2$
4. Find the concentration of the total $n(OH^-)$. *Remember to use the volume of the total solution*
5. Calculate pH as per previous pages

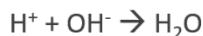
Question – Mixing bases

An environmental chemist is conducting research into highly alkaline solutions found in a drain near a concrete batching plant. The concentration of the solution from the drain was found to be 0.236 mol L^{-1} hydroxide ion and the runoff from the plant was found to have a concentration of hydroxide ion of 0.156 mol L^{-1} . The chemist mixed 200 mL of the solution from the drain with 300 mL of the runoff from the plant.

Calculate the pH of the resulting solution:

Calculating pH – Mixing acid + base

Sometimes you may have a question where a strong acid is mixed with a strong base. Acids and bases neutralise each other...



...but if **excess acid** is used then the resulting solution will be **acidic**, and if **excess base** is used then the resulting solution will be **basic**.

You may need to calculate the pH of such a solution after the acid and base have been mixed together.

Method:

1. Find $n(\text{H}^+)$
2. Find $n(\text{OH}^-)$
3. Figure out which substance is in excess. Calculate $n(\text{H}^+ \text{ excess})$ or $n(\text{OH}^- \text{ excess})$
4. Find the concentration of the excess ions. *Remember to use the volume of the total solution*
5. Calculate pH as per previous pages

Example: 100 mL of 1 mol L⁻¹ Ca(OH)₂ is added to 400 mL of 0.300 mol L⁻¹ HCl. What is the pH of the new solution?

1 $n(\text{Ca}(\text{OH})_2) = c \times V$
 = 1 × 0.1
 = 0.1 mol

$$n(\text{OH}^-) = 2 \times n(\text{Ca}(\text{OH})_2)$$
$$= 0.2 \text{ mol}$$

2 $n(\text{HCl}) = c \times V$
 = 0.300 × 0.400
 = 0.120 mol

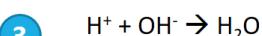
$$n(\text{H}^+) = 0.120 \text{ mol}$$

4 Total volume = 500 mL = 0.5 L

$$c(\text{OH}^- \text{ excess}) = n / V$$
$$= 0.080 / 0.500$$
$$= 0.16 \text{ mol L}^{-1}$$

5 $\text{pOH} = -\log[\text{OH}^-]$
 = $-\log[0.16]$
 = 0.796

$$\text{pH} = 14 - \text{pOH}$$
$$= 13.2$$



$$n(\text{OH}^- \text{ excess}) = n(\text{OH}^-) - n(\text{H}^+)$$
$$= 0.200 - 0.120$$
$$= 0.080 \text{ mol}$$

Questions – Mixed acid-base pH calculations

Calculate the pH when the following solutions are mixed:

- a) 200 mL of 0.1 mol L⁻¹ sodium hydroxide and 300 mL of 0.1 mol L⁻¹ hydrochloric acid

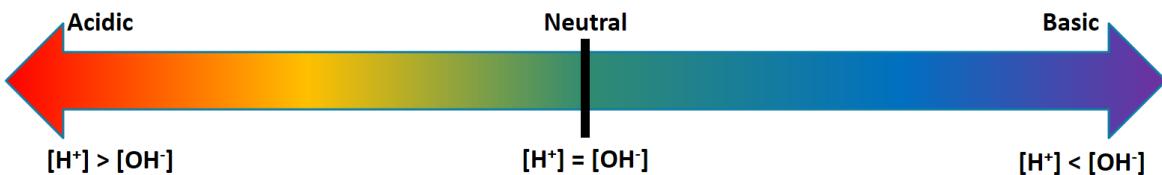
- b) 200 mL of 0.1 mol L⁻¹ sodium hydroxide and 300 mL of 0.1 mol L⁻¹ sulfuric acid

- c) When you suffer from an acid stomach, the fluid in your stomach can contain hydrochloric acid at a concentration of $2.50 \times 10^{-4} \text{ mol L}^{-1}$. To reduce the amount of acid in your stomach you are advised to take an antacid tablet that contains 5.00 mg of aluminium hydroxide and 5.00 mg of magnesium hydroxide. If at this time you have 1.50 L of fluid in your stomach, calculate the $[\text{H}^+]$, $[\text{OH}]$ and pH of your stomach after you have taken the tablet. Assume that the presence of the tablet does not change the volume of the solution in your stomach.

Acidic, Basic & Neutral

A **neutral solution** is defined as a solution with equal concentrations of H^+ and OH^- .

Even though pure water self-ionises to produce small amounts of $[\text{H}^+]$ and $[\text{OH}^-]$, the resulting solution will still be neutral because $[\text{H}^+] = [\text{OH}^-]$



pH of Neutral Solutions

We can calculate the pH of neutral water at 25 °C by combining two equations:

$$\text{Equation 1: } [\text{H}^+] = [\text{OH}^-]$$

This equation is true for any neutral solution

$$\text{Equation 2: } K_w = [\text{H}^+] [\text{OH}^-]$$

Equilibrium constant for water.

$$1 \times 10^{-14} = [\text{H}^+] [\text{OH}^-]$$

This is the value of K_w at 25 °C

Example: Show using calculations that the pH of neutral water at 25 °C is equal to 7.

$$K_w = [\text{H}^+] [\text{OH}^-]$$

The solution is neutral, ∴ $[\text{H}^+] = [\text{OH}^-]$

$$= [\text{H}^+]^2$$

At 25 °C, $K_w = 1.0 \times 10^{-14}$

$$1.0 \times 10^{-14} = [\text{H}^+]^2$$

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

$$= 1 \times 10^{-7}$$

$$\text{pH} = -\log[\text{H}^+]$$

$$= -\log[1 \times 10^{-7}]$$

$$= 7.00$$

Effect of temperature of pH

Like all equilibrium reactions, the self-ionisation of water is affected by **temperature**.

Investigate the table to the right. Is the self-ionisation of water an **exothermic** or **endothermic process**?

Explain your answer in terms of Le Châtelier's principle.

Temperature (°C)	K_w
0	0.11×10^{-14}
10	0.29×10^{-14}
20	0.68×10^{-14}
25	1.01×10^{-14}
30	1.47×10^{-14}
40	2.92×10^{-14}
50	5.48×10^{-14}
100	51.3×10^{-14}

You should be able to calculate the pH of neutral water at any temperature if given the value of K_w .

The working will be similar to that on the previous page

Q: Calculate the pH of water at the following temperatures. Use the K_w values from the table above.

a) Neutral water at 10 °C

b) Neutral water at 50 °C

Buffers

Buffers are solutions that **resists changes to pH**.

They are made of a **weak acid** and its **conjugate base**.

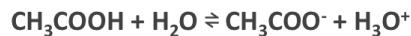
Examples:

- A mixture of **acetic acid** (CH_3COOH) and sodium **acetate** (NaCH_3COO)
- A mixture of **ammonia** (NH_3) and **ammonium** chloride (NH_4Cl)
- A mixture of **phosphoric acid** (H_3PO_4) and sodium **phosphate** (Na_3PO_4)

How buffers work

Consider a solution containing acetic acid and sodium acetate.

Inside the solution, the acetic acid and acetate would be in **equilibrium**.



Explain why pH would remain roughly constant when adding acid:

- When acid is added this would initially **increase** $[\text{H}_3\text{O}^+]$
- According to Le Châtelier's Principle the system will react to **oppose** this change.
- The reaction will shift to the left, **decreasing** $[\text{H}_3\text{O}^+]$ close to its original value
- As $[\text{H}_3\text{O}^+]$ does not change greatly, there is little change to pH

Use Le Châtelier's principle to explain why pH does not change when sodium hydroxide is added to a mixture of acetic acid and sodium acetate.

Use Le Châtelier's principle to explain why pH does not change when acid is added to a mixture of sodium hydrogencarbonate and sodium carbonate.

Explain why a mixture of nitric acid and sodium nitrate would **not** make a good buffer.

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Buffering capacity

Buffering capacity is a measure of how well a buffer can resist changes to pH.

A buffer with a **high buffering capacity** can have a **lot** of acid and base added before pH changes

A buffer with a **low buffering capacity** can only prevent changes for a small amount of acid or base.

Two factors affect buffering capacity:

➤ The relative amounts of the weak acid and its conjugate base.

Buffers work best with equimolar amounts. e.g. 0.5 mol CH₃COOH + 0.5 mol CH₃COO⁻ c.f. 0.5 mol + 0.1 mol

➤ The overall concentration of the buffer

High concentration of buffers are better. e.g. 2 mol CH₃COOH + 2 mol CH₃COO⁻ c.f. 0.5 mol + 0.5 mol

Water contains H₃O⁺ ions and OH⁻ ions due to the self-ionsation of water. 2 H₂O(l) ⇌ H₃O⁺(aq) + OH⁻(aq).

Adding acid or base to pure water would cause this equilibrium system to shift to the left, counteracting the imposed change. However, when acid or base are actually added to water the pH is observed to change rapidly. Explain why pure water does not have significant buffering capacity.

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EXPERIMENT: MAKING AND TESTING BUFFERS

Equipment:

- 4 beakers (100 mL)
- Measuring cylinder (50 mL)
- Glass stirring rod
- 2 plastic pipettes
- 1 mol L⁻¹ solutions of:
 - acetic acid solution (10 mL)
 - sodium acetate solution (40 mL)
- 0.1 mol L⁻¹ solutions of:
 - hydrochloric acid solution (30 mL)
 - sodium hydroxide solution (10 mL)
- Universal indicator

Method:

1. Place 10 mL of 1 mol L⁻¹ CH₃COOH into a clean dry 100 mL beaker. To this add 40 mL of 1 mol L⁻¹ CH₃COONa and mix thoroughly.
2. Transfer 25 mL of this solution to another clean dry 100 mL beaker. Label both these beakers as acetic acid-acetate buffer.
3. Add 5 drops of universal indicator to each of the beakers. Using the universal indicator chart, estimate and record the pH of the solution. Retain these for further testing.
4. Place 50 mL of distilled water into a 100 mL beaker then add 10 drops of universal indicator. Add 0.1 mol L⁻¹ HCl solution dropwise until the colour of the indicator is the same as in the buffer solution.
5. Transfer 25 mL of this solution into another clean dry beaker. Label both beakers as unbuffered acid solution
6. a) To one of the beakers containing acetic acid-acetate buffer add one drop of 0.1 mol L⁻¹ HCl, stir the solution then record the colour. Repeat this until you have added 40 drops. Estimate the pH after the 40 drops have been added.
b) Repeat this with one of the beakers containing unbuffered solution. Again, record any colour changes after the addition of each drop and estimate the pH after 40 drops have been added
7. Repeat procedure 6(a) and 6(b) with the remaining beaker of buffered and unbuffered solution, but this time use the 0.1 mol L⁻¹ NaOH solution

Processing of results:

Compare the behaviour of the buffered solutions with the unbuffered solutions when either acid or base was added.

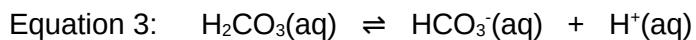
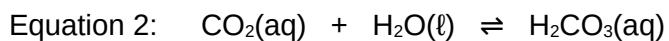
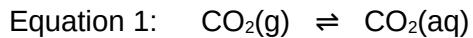
Explain the behaviour of the acetic acid-acetate buffer as...

- (i) H⁺ ion is added:
-
-
-
- (ii) OH⁻ ion is added:
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-
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Effect of Carbon Dioxide on our Oceans

What prior knowledge do you already have about the levels of carbon dioxide in the atmosphere?

Study the following three equations below which have been proposed to occur in the environment:

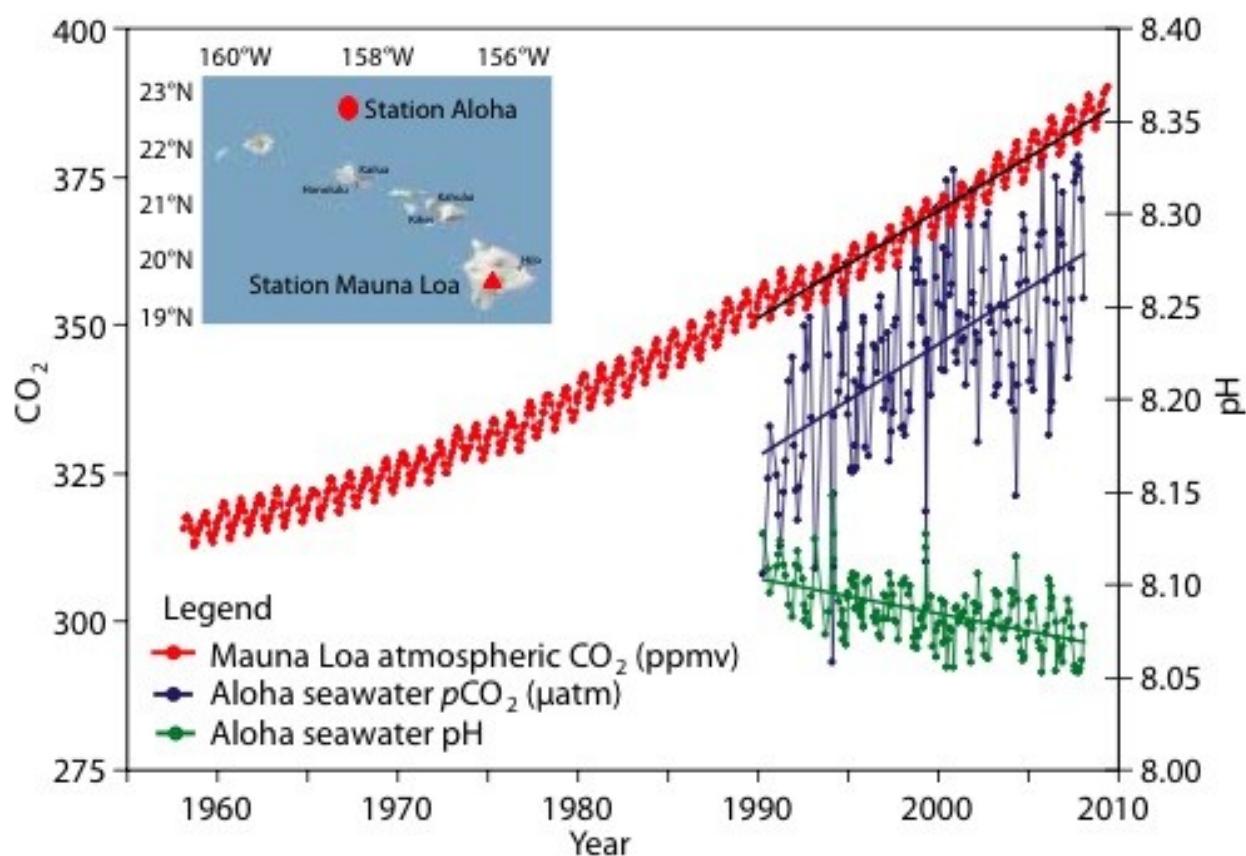


(Recall that H_2CO_3 is carbonic acid, a weak acid that partially ionises in solution to form H^+)

Using your knowledge of **chemical equilibrium**, predict how increasing levels of atmospheric CO_2 will affect the three equilibrium reactions.

Write a hypothesis relating atmospheric CO_2 levels and pH of surrounding water in contact with the CO_2 .

Scientists have collected empirical data which can be used to test your hypothesis on a global scale. The following graph shows data from two research stations in Hawaii. One station has measured atmospheric CO₂ since the 1950s. The other station has measured dissolved CO₂ and seawater pH since the 1990s.



Does the empirical data shown in the graph support your hypothesis? Explain.

What additional data do you think would be important in further testing this hypothesis?

Over the past 300 million years the average pH of the ocean has been approximately 8.2, which is slightly basic. Today, it is around pH 8.1. Although this is only a difference of 0.1 pH units, it represents a 25% increase acidity (*pH is a logarithmic scale; 1 pH unit corresponds to a 10x increase in H⁺ concentration*).

- What could be the effects of ocean acidification? Brainstorm with a partner.
- Select one of the potential effects from your brainstorm. Consider how this could be investigated by scientists. Write a valid **hypothesis** for a relevant scientific investigation and outline the sort of data that would need to be collected.

Hypothesis:

What data would need to be collected?

What would be the challenges associated with this investigation?

'Acid Ocean' documentary: <http://webplayer.clickview.com.au/ev.aspx?p=7Vd5q&ssi=0&a=true>

List the effects of ocean acidification described in the video:

Starting at ~8 minutes in, the documentary focuses on the field work conducted by scientists near Papua New Guinea. Describe why this site provides a unique opportunity for conducting studies related to ocean acidification.

Read the following syllabus dot point carefully. Each sentence in here indicates **examinable content**.

Levels of carbon dioxide in the atmosphere are rising and have a significant impact on global systems, including surface temperatures. The increasing level of carbon dioxide in the atmosphere causes more carbon dioxide to dissolve in the ocean producing carbonic acid and leading to increased ocean acidity. This is predicted to have a range of negative consequences for marine ecosystems such as coral reefs. Calcification is the process which results in the formation of calcium carbonate structures in marine organisms. Acidification shifts the equilibrium of carbonate chemistry in seawater, decreasing the rate and amount of calcification among a wide range of marine organisms. The United Nations Kyoto Protocol and the Intergovernmental Panel on Climate Change aim to secure a global commitment to reducing greenhouse gas emissions over the next few decades.

You have to be able to answer questions related to **any** and **all** of this content. One strategy you should use with the syllabus is to figure out for each sentence, what questions can they ask. The following tables will guide you through that process for each sentence of the ocean acidification dot point. Use your textbook (**page 20-22**) and/or computer research to figure out how you could answer each question.

Levels of carbon dioxide in the atmosphere are rising and have a significant impact on global systems, including surface temperatures.

<p>Question:</p> <p>Explain the changing trends in atmospheric carbon dioxide levels. (3 marks)</p>	<p>Answer:</p> <ul style="list-style-type: none">Atmospheric carbon dioxide levels have been increasing due to human activity.Burning of fossil fuels results in the emission of carbon dioxide which is released into the atmosphere.Deforestation and land clearing removes plants from the environment which would otherwise remove carbon dioxide from the air through photosynthesis.
<p>Question:</p> <p>Explain how the presence of carbon dioxide in the atmosphere has affected surface temperatures. (4 marks)</p>	<p>Answer:</p> <ul style="list-style-type: none">Atmospheric carbon dioxide acts as a greenhouse gas.The Earth emits heat radiation out into spaceCO_2 molecules in atmosphere absorb this radiation and re-emit it towards EarthThis causes a build up of heat on Earth, therefore increasing surface temperatures

The increasing level of carbon dioxide in the atmosphere causes more carbon dioxide to dissolve in the ocean producing carbonic acid and leading to increased ocean acidity.

Question:	Answer:
<p>Explain (using collision theory) why increasing concentrations of atmospheric CO₂ can lead to ocean acidification.</p> <p>Include relevant chemical equations as part of your answer. (7 marks)</p> <p><i>(The marks here are a suggested guide. 7 marks indicates seven ideas, including the three chemical equations.)</i></p>	

This is predicted to have a range of negative consequences for marine ecosystems such as coral reefs.

Question:

List a range of negative consequences of ocean acidification for marine ecosystems. (4 marks)

Answer:

Calcification is the process which results in the formation of calcium carbonate structures in marine organisms.

Describe the process of calcification, using an appropriate chemical equation. (2 marks)

Answer:

Acidification shifts the equilibrium of carbonate chemistry in seawater, decreasing the rate and amount of calcification among a wide range of marine organisms.

Question:

Explain how ocean acidification has affected the rate and amount of calcification in marine organisms. (5 marks)

Answer:

The United Nations Kyoto Protocol and the Intergovernmental Panel on Climate Change aim to secure a global commitment to reducing greenhouse gas emissions over the next few decades.

Question: What is the United Nations Kyoto Protocol? (2 marks)	Answer:
Question: What is the role of the Intergovernmental Panel on Climate Change? Who are they and what do they do? (2 marks)	Answer:

HOMEWORK

Source	Set/ Chapter	Topics	Questions to complete													
Essential Chemistry	Set 3	Review: Acid-Base Properties	1	2	3	4										
	Set 4	Models of acids and bases	1	2	3											
	Set 4	Bronsted-Lowry Theory	4	5	6	7	8	9	10	11						
	Set 4	Acid-Base equilibrium, K_a	12	13	14	15										
	Set 4	Hydrolysis of salts	16	17	18	19	20	21	22							
	Set 5	K_w and pH	1	2	3	4	5	6	7	8	9	10	11			
	Set 6	Buffers	1	2	3	4	5	6								
Past Exam Questions	Section 1	Bronsted-Lowry theory, conjugate acids and bases	[MC]	Pg:	2	3	4									
			[SA]	Pg:	5	6	7									
	Section 2	Polyprotic acids, acid strength, hydrolysis of salts	[MC]	Pg:	8	9	10	11								
			[SA]	Pg:	12	13	14	15	16							
	Section 3	Self-ionisation of water, K_w , pH calculations	[MC]	Pg:	17	18	19	20								
			[SA]	Pg:	21	22	23	24	25	26						
	Section 4	Equilibrium systems, acid dissociation constants, buffers	[MC]	Pg:	27	28										
			[SA]	Pg:	29	30	31	32	33	34	35	36	37			
	Section 5	Ocean acidification, historical models of acids and bases	[MC]	Pg:	38											
			[SA]	Pg:	39	40										