

CLASS WORKSHEET 5.1 - ACIDS, BASES AND SALTS

1) Review of Cations, Anions and Molecules

CATIONS		ANIONS		MOLECULES	
ammonium	NH_4^+	bromide	Br^-	ammonia	NH_3
aluminium	Al^{3+}	carbonate	CO_3^{2-}	carbon dioxide	CO_2
barium	Ba^{2+}	chloride	Cl^-	water	H_2O
calcium	Ca^{2+}	fluoride	F^-		
hydrogen	H^+	hydroxide	OH^-		
lithium	Li^+	iodide	I^-		
magnesium	Mg^{2+}	nitrate	NO_3^-		
potassium	K^+	oxide	O^{2-}		
sodium	Na^+	sulfate	SO_4^{2-}		
strontium	Sr^{2+}				

2) Formulae of Ionic Compounds

Name	Formula	Name	Formula
sodium oxide	Na_2O	hydrogen chloride	HCl
calcium hydroxide	$\text{Ca}(\text{OH})_2$	rubidium hydroxide	RbOH
ammonium nitrate	NH_4NO_3	magnesium carbonate	MgCO_3
potassium carbonate	K_2CO_3	calcium nitrate	$\text{Ca}(\text{NO}_3)_2$
strontium sulfate	SrSO_4	hydrogen sulfate	H_2SO_4
ammonium sulfate	$(\text{NH}_4)_2\text{SO}_4$	ammonium chloride	NH_4Cl
hydrogen nitrate	HNO_3	caesium bromide	CsBr
potassium sulfate	K_2SO_4	barium sulphate	BaSO_4
magnesium oxide	MgO	strontium nitrate	$\text{Sr}(\text{NO}_3)_2$

3) Definitions of acids, bases and salts

(i) Acids

Eg $\text{HCl}(\text{g}) \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

A mixture of HCl and water contains H^+ ions and is known as **hydrochloric acid**

Eg $\text{H}_2\text{SO}_4(\text{l}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

A mixture of H_2SO_4 and water contains H^+ ions and is known as **sulfuric acid**

Eg $\text{HNO}_3(\text{l}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$

A mixture of HNO_3 and water contains H^+ ions and is known as **nitric acid**

H^+ ions can also be referred to as **protons**.

Acids can therefore be described as **proton donors**.

(ii) bases and alkalis

Bases can therefore be described as **proton acceptors**.

Hydroxide ions can accept H^+ ions to form H_2O : $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

Oxide ions can accept H^+ ions to form H_2O : $2\text{H}^+ + \text{O}^{2-} \rightarrow \text{H}_2\text{O}$

Carbonate ions can accept H^+ ions to form CO_2 and H_2O : $2\text{H}^+ + \text{CO}_3^{2-} \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

Ammonia can accept H^+ ions to form NH_4^+ ions: $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$

A substance which dissolves in water to give a solution containing OH⁻ ions is called an **alkali**. An alkali is therefore a **soluble base**.

4) Classifying substances as acids, bases and salts

ACIDS	BASES	SALTS
HNO ₃	Na ₂ O	NH ₄ NO ₃
HCl	Ca(OH) ₂	SrSO ₄
H ₂ SO ₄	K ₂ CO ₃	(NH ₄) ₂ SO ₄
	MgO	K ₂ SO ₄
	MgCO ₃	Ca(NO ₃) ₂
	RbOH	NH ₄ Cl
		CsBr
		BaSO ₄
		Sr(NO ₃) ₂

5) Neutralization Reactions

Acid + metal hydroxide → metal salt + water

Eg hydrochloric acid + sodium hydroxide → sodium chloride + water

Symbol equation: $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

Acid + metal oxide → metal salt + water

Eg nitric acid + calcium oxide → calcium nitrate + water

Symbol equation: $2\text{HNO}_3 + \text{CaO} \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$

Acid + metal carbonate → metal salt + carbon dioxide + water

Eg sulphuric acid + potassium carbonate → potassium sulphate + carbon dioxide + water

Symbol equation: $\text{H}_2\text{SO}_4 + \text{K}_2\text{CO}_3 \rightarrow \text{K}_2\text{SO}_4 + \text{CO}_2 + \text{H}_2\text{O}$

Acid + ammonia → ammonium salt

Eg hydrochloric acid + ammonia → ammonium chloride

Symbol equation: $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$

a) Nitric acid with potassium hydroxide solution

Equation: $\text{HNO}_3(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{KNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$

Name of salt: **potassium nitrate**

b) Sulfuric acid with sodium hydroxide solution

Equation: $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$

Name of salt: **potassium sulfate**

c) Hydrochloric acid with calcium hydroxide powder

Equation: $2\text{HCl}(\text{aq}) + \text{Ca}(\text{OH})_2(\text{s}) \rightarrow \text{CaCl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$

Name of salt: **calcium chloride**

d) Nitric acid with calcium oxide powder



Name of salt: calcium nitrate

e) Hydrochloric acid with barium oxide powder



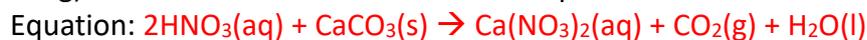
Name of salt: barium chloride

f) Sulfuric acid with magnesium oxide powder



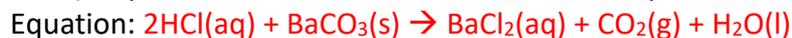
Name of salt: magnesium sulfate

g) Nitric acid with calcium carbonate powder



Name of salt: calcium nitrate

h) Hydrochloric acid with barium carbonate powder



Name of salt: barium chloride

i) Sulphuric acid with sodium carbonate solution



Name of salt: sodium sulfate

j) Nitric acid with ammonia solution



Name of salt: ammonium nitrate

k) Sulfuric acid with ammonia solution



Name of salt: ammonium sulfate

l) Hydrochloric acid with ammonia solution



Name of salt: ammonium chloride

OBSERVING NEUTRALISATION REACTIONS

When an acid reacts with a solid metal hydroxide or oxide, the solid **dissolves**.

When an acid reacts with a solid metal carbonate, the solid **dissolves** and **effervescence** will also be observed.

When an acid reacts with an aqueous metal carbonate, **effervescence** will be observed

5.2 HONORS CLASS WORKSHEET – ACIDITY, ALKALINITY AND THE PH SCALE

1) Acidity and Alkalinity

The ion which makes solutions acidic is H^+

The ion which makes solutions alkaline is OH^-

Water dissociates very slightly to produce H^+ and OH^- ions. Equation: $H_2O \rightleftharpoons H^+ + OH^-$

In pure water, the concentration of H^+ and OH^- is around 1×10^{-7} mol/L

Any solution which contains equal concentrations of H^+ and OH^- ions is said to be **neutral**

In solutions which contain acids, how will the concentrations of H^+ and OH^- compare to those in pure water, and therefore to each other?

$[H^+] > 1 \times 10^{-7}$ mol/L, $[OH^-] < 1 \times 10^{-7}$ mol/L, $[H^+] > [OH^-]$

In solutions which contain alkalis, how will the concentrations of H^+ and OH^- compare to those in pure water, and therefore to each other?

$[H^+] < 1 \times 10^{-7}$ mol/L, $[OH^-] > 1 \times 10^{-7}$ mol/L, $[H^+] < [OH^-]$

Concentration of H^+ ions (mol/L)	Concentration of OH^- ions (mol/L)	Type of solution
0.1 (1×10^{-1})	1×10^{-13}	acidic
0.001 (1×10^{-3})	1×10^{-11}	acidic
1×10^{-5}	1×10^{-9}	acidic
1×10^{-7}	1×10^{-7}	neutral
1×10^{-9}	1×10^{-5}	alkaline
1×10^{-11}	1×10^{-3} (0.001)	alkaline
1×10^{-13}	1×10^{-1} (0.1)	alkaline

2) The pH scale

The pH of a solution is defined as **the negative logarithm of the hydrogen ion concentration** (pH stands for power of hydrogen)

pH is a logarithmic scale. What does this mean? **An change by a factor of 10 in the quantity causes a change of 1 unit in the scale**

- if the H^+ concentration is 0.1 (ie 1×10^{-1}) mol/L, the pH of the solution is **1**
- if the H^+ concentration is 0.001 (ie 1×10^{-3}) mol/L, the pH of the solution is **3**
- if the H^+ concentration is 1×10^{-7} mol/L, the pH of the solution is **7**
- if the H^+ concentration is 1×10^{-11} mol/L, the pH of the solution is **11**
- if the H^+ concentration is 1×10^{-13} mol/L, the pH of the solution is **13**

What does a low pH tell you about the solution? **acidic**

What does a high pH tell you about the solution? **alkaline**

The relationship between pH, acidity and alkalinity is summarised in the table below:

pH	-1	1	3	5	7	9	11	13	15
Acidity	strongly acidic		weakly acidic		neutral		weakly alkaline		strongly alkaline
$[H^+]$	very high		quite high		medium		quite low		very low
$[OH^-]$	very low		quite low		medium		quite high		very high

Examples of the pH of common solutions are:

solution	pH	Solution	pH	solution	pH
1 mol/L HCl	0	lemon juice	2/3	vinegar	3
orange juice	3/4	pure water	7	household bleach	13
1 mol/L NaOH	14				

5.3 HONORS CLASS WORKSHEET – STRONG AND WEAK ACIDS

1) strong and weak acids

Strong Acid: fully dissociates into H^+ ions in water

Example: HCl Equation: $HCl \rightarrow H^+ + Cl^-$

Weak Acid: slightly dissociates into H^+ ions in water

Example: acetic acid – $HC_2H_3O_2$ Equation: $HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$

2) strong and weak bases

Strong Base: fully dissociates into OH^- ions in water

Example: NaOH Equation: $NaOH \rightarrow Na^+ + OH^-$

Weak Base: slightly dissociates into OH^- ions in water

Example: NH_3 Equation: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Example CO_3^{2-} Equation: $CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$

Eg $Ca(OH)_2$ Equation: $Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$

3) Neutralizing strong and weak acids

0.01 moles of a strong acid (HCl) will require 0.01 moles of OH^- to neutralise it

0.01 moles of a weak acid ($HC_2H_3O_2$) will require 0.01 moles of OH^- to neutralise it

Reason: Equation 1: $HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$

Equation 2: $H^+ + OH^- \rightarrow H_2O$

As the H^+ ions react with the OH^- , the weak acid dissociates more to replace them; this continues until the weak acid is fully dissociate

4) Differences between strong and weak acids (and bases)

(i) Enthalpy of neutralization

Equation for neutralization of HCl/ HNO_3 : $H^+ + OH^- \rightarrow H_2O$ $\Delta H = -57 \text{ kJ/mol}$

Equation for neutralization of $HC_2H_3O_2$:

1. $HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$ $\Delta H = +5 \text{ kJ/mol}$

2. $H^+ + OH^- \rightarrow H_2O$ $\Delta H = -57 \text{ kJ/mol}$

Overall: $\Delta H = -52 \text{ kJ/mol}$

(ii) pH

0.100 mol/L HCl is 100% dissociated so H^+ concentration = 0.100 mol/L and pH = 1

0.100 mol/L $HC_2H_3O_2$ is 1% dissociated so H^+ concentration = 0.001 mol/L and pH = 3

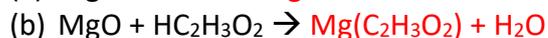
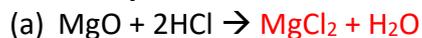
0.100 mol/L NaOH is 100% dissociated so OH^- concentration = 0.100 mol/L and pH = 13

0.100 mol/L NH_3 is 1% dissociated so OH^- concentration = 0.001 mol/L and pH = 11

(iii) Conductivity

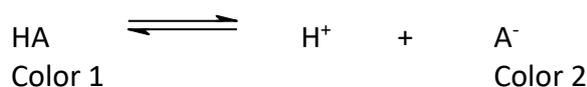
Which solution will have the higher conductivity: 0.1 mol/L HCl or 0.1 mol/L HC₂H₃O₂? Why?
0.1 mol/L HCl will have a higher conductivity than 0.1 mol/L HC₂H₃O₂ because the acid is fully dissociated into its ions, so the concentration of ions in solution is much higher

(iv) Reactivity



Which reaction will be faster, and why? The reaction between MgO and HCl will be much faster because HCl is a strong acid so has a much higher concentration of H⁺ ions

5.4 HONORS CLASS WORKSHEET – ACID-BASE INDICATORS



State and explain which color the indicator will show at high pH

Color 2 as a high pH means a low [H⁺] which will push the equilibrium to the right

State and explain which color the indicator will show at low pH

Color 1 as a low pH means a high [H⁺] which will push the equilibrium to the left

The pH at which the indicator changes color is called the end point of the indicator.

It varies from indicator to indicator. It depends on the strength of the acid, which affects its tendency to dissociate

Indicator	Color 1	Color 2	pH at which color changes	color during transitional pH range
methyl orange	pink	yellow	3.1 – 4.4	orange
methyl red	red	yellow	4.4 – 6.3	orange
bromothymol blue	yellow	blue	6.0 – 7.7	green
phenolphthalein	colorless	purple	8.3 – 10.0	pink/purple

pH	methyl orange	methyl red	bromothymol blue	phenolphthalein	Mixture of all indicators
2.0	pink	red	yellow	colorless	red/orange
3.5	orange	red	yellow	colorless	orange
5.0	yellow	orange	yellow	colorless	yellow
6.5	yellow	yellow	green	colorless	turquoise
8.0	yellow	yellow	blue	colorless	green
9.5	yellow	yellow	blue	pink	blue
11.0	yellow	yellow	blue	Purple	blue/purple

5.5 HONORS CLASS WORKSHEET – TITRATIONS

Quantitative Analysis: determination of how much of a substance is present

Volumetric Analysis: determination of how much of a substance is present by measuring volumes

Titration: precise determination of the volume of one liquid required to completely react with a fixed volume of another

Steps involved in carrying out a titration:

- use a pipette to measure out a fixed volume of alkali into a conical flask
- add two drops of indicator to the flask
- fill a burette with acid and record the initial volume of acid in the burette
- allow the acid from the burette to run slowly into the conical flask
- until the indicator changes color
- note the final volume of acid in the burette
- subtract the initial volume from the final volume to determine the volume of acid used

Why should the titration be carried out three times?

- 2nd result determines whether or not the two results are consistent
- 3rd result states which of the 1st two titrations is more reliable, if the first two are not consistent

The point during a titration at which the acid and the alkali have neutralised each other exactly is called the **equivalence point**.

Type of titration	pH change at equivalence point	best indicator
strong acid - strong base	3.0 – 11.0	any
weak acid - strong base	7.0 – 11.0	phenolphthalein
strong acid - weak base	3.0 – 7.0	methyl red

Question: In a titration 28.3 mL of a 0.10 mol/L solution of H₂SO₄ was required to change the color of the indicator in 25 mL of a solution of NaOH. What was the molarity of the NaOH solution?

Equation:	$\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$	
	acid	Base
volume (mL)	28.3	25
Moles	0.00283	0.00566
molarity (mol/L)	0.1	0.226

- (a) 18.4 mL of HCl was required to neutralise 25 mL of 0.1 mol/L NaOH. Deduce the molarity of the HCl.

Equation:	$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$	
	acid	Base
volume (mL)	18.4	25
moles	0.0025	0.0025
molarity (mol/L)	0.136	0.1

- (b) 13.9 mL of acetic acid (HC₂H₃O₂) was required to neutralise 25 mL of 0.1 mol/L NaOH. Deduce the molarity of the HC₂H₃O₂.

Equation:	$\text{HC}_2\text{H}_3\text{O}_2 + \text{NaOH} \rightarrow \text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O}$	
	acid	Base
volume (mL)	13.9	25
moles	0.0025	0.0025
molarity (mol/L)	0.18	0.1

- (c) 25.0 mL of a solution of Na_2CO_3 was titrated against 0.1 mol/L HCl and 24.5 mL of the acid were required. Calculate the molarity of the Na_2CO_3 solution.

Equation:	$2\text{HCl} + \text{Na}_2\text{CO}_3 \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$	
	Acid	Base
volume (mL)	24.5	25
Moles	0.00245	0.001225
molarity (mol/L)	0.1	0.049

- (d) Sodium carbonate exists in hydrated form, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, in the solid state. 0.35 g of a sodium carbonate sample was dissolved in water. The resulting solution was titrated against 0.1 mol/L HCl and 24.5 mL of the acid were required. Calculate the value of x given the equation: $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{CO}_2 + \text{H}_2\text{O}$

$$\text{moles of HCl} = 24.5/1000 \times 0.1 = 0.00245$$

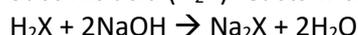
$$\text{moles of Na}_2\text{CO}_3 = 0.001225$$

$$\text{molar mass of Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = 0.35/0.001225 = 286$$

$$\text{Na}_2\text{CO}_3 = 106 \text{ so } x\text{H}_2\text{O} = 18x = (286 - 106) = 180$$

$$\text{So } x = 180/18 = 10$$

- (e) Succinic acid (H_2X) reacts with dilute sodium hydroxide as follows:



2.00 g of succinic acid were dissolved in water and used to prepare a 250 mL solution. This solution was placed in a burette and 18.4 mL was required to neutralise 25 mL of 0.1 mol/L NaOH. Deduce the molar mass of succinic acid.

- (i) Calculate the mass concentration of the succinic acid solution

$$2.00/(250/1000) = 8.00 \text{ g/L}$$

- (ii) Use the titration result to deduce the molarity of the succinic acid solution

$$\text{Moles of NaOH} = 25/1000 \times 0.1 = 0.0025, \text{ so moles of H}_2\text{X} = 0.00125$$

$$\text{Volume of H}_2\text{X} = 18.4 \text{ mL} = 0.0184 \text{ L}, \text{ so molarity} = 0.00125/0.0184 = 0.0679 \text{ mol/L}$$

- (iii) Hence calculate the molar mass of succinic acid

$$\text{Molar mass} = \text{mass concentration} / \text{molarity} = 8.00/0.0679 = 118 \text{ g/mol}$$

- (f) Oxalic acid ($\text{H}_2\text{C}_2\text{O}_4 \cdot x\text{H}_2\text{O}$) reacts with NaOH in a 1:2 ratio. 1.85 g of oxalic acid was dissolved in water and used to prepare 250 mL of oxalic solution. This solution was placed in a burette and 21.3 mL were required to neutralise 25 mL of 0.1 mol/L NaOH.

- (i) Calculate the mass concentration of the oxalic acid solution

$$1.85/(250/1000) = 7.40 \text{ g/L}$$

- (ii) Use the titration result to deduce the molarity of the oxalic acid solution

$$\text{Moles of NaOH} = 25/1000 \times 0.1 = 0.0025, \text{ so moles of H}_2\text{C}_2\text{O}_4 \cdot x\text{H}_2\text{O} = 0.00125$$

$$\text{Volume of H}_2\text{X} = 21.3 \text{ mL} = 0.0213 \text{ L}, \text{ so molarity} = 0.00125/0.0213 = 0.0586 \text{ mol/L}$$

- (iii) Hence calculate the molar mass of oxalic acid

$$\text{Molar mass} = \text{mass concentration} / \text{molarity} = 7.40/0.0586 = 126 \text{ g/mol}$$

- (iv) Hence calculate the value of x

$$\text{H}_2\text{C}_2\text{O}_4 = 90, \text{ so } 90 + 18x = 126, \text{ so } 18x = 36, \text{ so } x = 36/18 = 2$$