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	$Ca(OH)_2$	rubidium hydroxide	$RbOH$
ammonium nitrate	NH_4NO_3	magnesium carbonate	$MgCO_3$
potassium carbonate	K_2CO_3	calcium nitrate	$Ca(NO_3)_2$
strontium sulfate	$SrSO_4$	hydrogen sulfate	H_2SO_4
ammonium sulfate	$(NH_4)_2SO_4$	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	$Sr(NO_3)_2$

3) Definitions of acids, bases and salts

(i) Acids

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

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

Eg $H_2SO_4(l) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

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

Eg $HNO_3(l) \rightarrow H^+(aq) + NO_3^-(aq)$

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

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

Carbonate ions can accept H^+ ions to form CO_2 and H_2O: $2H^+ + CO_3^{2-} \rightarrow CO_2 + H_2O$

Ammonia can accept H^+ ions to form NH_4^+ ions: $NH_3 + H^+ \rightarrow NH_4^+$

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_2O	NH_4NO_3
HCl	$Ca(OH)_2$	$SrSO_4$
H_2SO_4	K_2CO_3	$(NH_4)_2SO_4$
	MgO	K_2SO_4
	$MgCO_3$	$Ca(NO_3)_2$
	$RbOH$	NH_4Cl
		$CsBr$
		$BaSO_4$
		$Sr(NO_3)_2$

5) Neutralization Reactions

Acid + metal hydroxide → metal salt + water

Eg hydrochloric acid + sodium hydroxide → sodium chloride + water

Symbol equation: $HCl + NaOH \rightarrow NaCl + H_2O$

Acid + metal oxide → metal salt + water

Eg nitric acid + calcium oxide → calcium nitrate + water

Symbol equation: $2HNO_3 + CaO \rightarrow Ca(NO_3)_2 + H_2O$

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

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

Symbol equation: $H_2SO_4 + K_2CO_3 \rightarrow K_2SO_4 + CO_2 + H_2O$

Acid + ammonia → ammonium salt

Eg hydrochloric acid + ammonia → ammonium chloride

Symbol equation: $HCl + NH_3 \rightarrow NH_4Cl$

a) Nitric acid with potassium hydroxide solution

Equation: $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

Name of salt: potassium nitrate

b) Sulfuric acid with sodium hydroxide solution

Equation: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$

Name of salt: potassium sulfate

c) Hydrochloric acid with calcium hydroxide powder

Equation: $2HCl(aq) + Ca(OH)_2(s) \rightarrow CaCl_2(aq) + 2H_2O(l)$

Name of salt: calcium chloride

d) Nitric acid with calcium oxide powder
Equation: $2HNO_3(aq) + CaO(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$
Name of salt: calcium nitrate

e) Hydrochloric acid with barium oxide powder
Equation: $2HCl(aq) + BaO(s) \rightarrow BaCl_2(aq) + H_2O(l)$
Name of salt: barium chloride

f) Sulfuric acid with magnesium oxide powder
Equation: $H_2SO_4(aq) + BaO(s) \rightarrow MgSO_4(aq) + H_2O(l)$
Name of salt: magnesium sulfate

g) Nitric acid with calcium carbonate powder
Equation: $2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$
Name of salt: calcium nitrate

h) Hydrochloric acid with barium carbonate powder
Equation: $2HCl(aq) + BaCO_3(s) \rightarrow BaCl_2(aq) + CO_2(g) + H_2O(l)$
Name of salt: barium chloride

i) Sulphuric acid with sodium carbonate solution
Equation: $H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$
Name of salt: sodium sulfate

j) Nitric acid with ammonia solution
Equation: $HNO_3(aq) + NH_3(aq) \rightarrow NH_4NO_3(aq)$
Name of salt: ammonium nitrate

k) Sulfuric acid with ammonia solution
Equation: $H_2SO_4(aq) + 2NH_3(aq) \rightarrow (NH_4)_2SO_4(aq)$
Name of salt: ammonium sulfate

l) Hydrochloric acid with ammonia solution
Equation: $HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$
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: $H_2O \rightleftarrows H^+ + OH^-$
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:

<table>
<tr><td>pH</td><td>-1</td><td>1</td><td>3</td><td>5</td><td>7</td><td>9</td><td>11</td><td>13</td><td>15</td></tr>
<tr><td>Acidity</td><td colspan="2">strongly acidic</td><td colspan="2">weakly acidic</td><td>neutral</td><td colspan="2">weakly alkaline</td><td colspan="2">strongly alkaline</td></tr>
<tr><td>[H⁺]</td><td colspan="2">very high</td><td colspan="2">quite high</td><td>medium</td><td colspan="2">quite low</td><td colspan="2">very low</td></tr>
<tr><td>[OH⁻]</td><td colspan="2">very low</td><td colspan="2">quite low</td><td>medium</td><td colspan="2">quite high</td><td colspan="2">very high</td></tr>
</table>

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: $HC_2H_3O_2 \rightleftarrows H^+ + C_2H_3O_2^-$

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: $NH_3 + H_2O \rightleftarrows NH_4^+ + OH^-$
Example CO_3^{2-} Equation: $CO_3^{2-} + H_2O \rightleftarrows HCO_3^- + OH^-$
Eg $Ca(OH)_2$ Equation: $Ca(OH)_2 \rightleftarrows Ca^{2+} + 2OH^-$

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: $HC_2H_3O_2 \rightleftarrows H^+ + C_2H_3O_2^-$
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: $H^+ + OH^- \rightarrow H_2O$ $\Delta H = -57$ kJ/mol
Equation for neutralization of $HC_2H_3O_2$:

1. $HC_2H_3O_2 \rightleftarrows H^+ + C_2H_3O_2^-$ $\Delta H = +5$ kJ/mol
2. $H^+ + OH^- \rightarrow H_2O$ $\Delta H = -57$ kJ/mol

Overall: $\Delta H = -52$ 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 $HC_2H_3O_2$? Why?
0.1 mol/L HCl will have a higher conductivity than 0.1 mol/L $HC_2H_3O_2$ because the acid is fully dissociated into its ions, so the concentration of ions in solution is much higher

(iv) Reactivity

(a) $MgO + 2HCl \rightarrow MgCl_2 + H_2O$

(b) $MgO + HC_2H_3O_2 \rightarrow Mg(C_2H_3O_2) + H_2O$

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

$$\underset{\text{Color 1}}{HA} \rightleftharpoons H^+ + \underset{\text{Color 2}}{A^-}$$

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 H_2SO_4 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:	$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$	
	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:	$HCl + NaOH \rightarrow NaCl + H_2O$	
	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 ($HC_2H_3O_2$) was required to neutralise 25 mL of 0.1 mol/L NaOH. Deduce the molarity of the $HC_2H_3O_2$.

Equation:	$HC_2H_3O_2 + NaOH \rightarrow NaC_2H_3O_2 + H_2O$	
	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:	$2HCl + Na_2CO_3 \rightarrow 2NaCl + CO_2 + H_2O$	
	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, $Na_2CO_3.xH_2O$, 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: $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$

moles of HCl = 24.5/1000 x 0.1 = 0.00245

moles of Na_2CO_3 = 0.001225

molar mass of $Na_2CO_3.xH_2O$ = 0.35/0.001225 = 286

Na_2CO_3 = 106 so xH_2O = 18x = (286 – 106) = 180

So x = 180/18 = 10

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

$H_2X + 2NaOH \rightarrow Na_2X + 2H_2O$

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 g/L

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

Moles of NaOH = 25/1000 x 0.1 = 0.0025, so moles of H_2X = 000125

Volume of H_2X = 18.4 mL = 0.0184 mL, so molarity = 0.00125/0.0184 = 0.0679 mol/L

(iii) Hence calculate the molar mass of succinic acid

Molar mass = mass concentration / molarity = 8.00/0.0679 = 118 g/mol

(f) Oxalic acid ($H_2C_2O_4.xH_2O$) 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 g/L

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

Moles of NaOH = 25/1000 x 0.1 = 0.0025, so moles of $H_2C_2O_4.xH_2O$ = 000125

Volume of H_2X = 21.3 mL = 0.0213 mL, so molarity = 0.00125/0.0213 = 0.0586 mol/L

(iii) Hence calculate the molar mass of oxalic acid

Molar mass = mass concentration / molarity = 7.40/0.0586 = 126 g/mol

(iv) Hence calculate the value of x

$H_2C_2O_4$ = 90, so 90 + 18x = 126, so 18x = 36, so x = 36/18 = 2