

ACIDS and BASES

Acids are chemicals when dissolved in water that produce _____ (_____).

Bases are chemicals that react with acids to produce a _____ and _____.

Alkalis are soluble bases. Examples of alkalis are _____.

Alkalis are chemicals when dissolved in water that produce _____ (_____).

Salts are ionic compounds produced when an _____ reacts with a _____.

PROPERTIES OF ACIDS

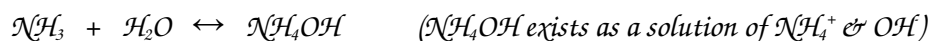
1. Litmus turns _____.
2. Universal Indicator turns _____ or _____ or _____.
3. Metals - react with acids to produce _____ and _____.
4. Bases - react with acids to produce _____ and _____.
5. Carbonates - react with acids to produce _____, _____ and _____.
6. Taste - sour. eg vinegar contains _____ acid.
lemon contains _____ acid.
soda water contains _____ acid.

PROPERTIES OF BASES

1. Litmus turns _____.
2. Universal Indicator turns _____ or _____ or _____.
3. Acids - react with bases to produce _____ and _____.
4. Ammonium salts - react with bases to produce a salt, ammonia and water. Therefore ammonium salts are weakly _____.

AMMONIA

Ammonia is a pungent smelling gas. It is soluble in water to produce a solution, which is called ammonia solution or ammonium hydroxide.



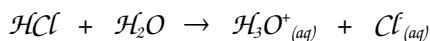
DESCRIBING ACIDS AND BASES

The following terms are used to describe acids and bases.

1) WEAK AND STRONG

Acids and bases ionise or dissociate (form ions in water). If most of the chemical ionises it is called a strong acid or a strong base. If only a little of the chemical ionises it is called a weak acid or base.

eg 1) HCl in water has nearly a 100% ionisation.



_____ (aq) is called a hydronium ion and is formed by joining a _____ and _____.

eg 2) NH_3 is a weak base because only 1% of NH_3 molecules react to form ions.
 $\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{NH}_4^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

Strong Acids $\text{HNO}_3, \text{HCl}, \text{HBr}, \text{HI}, \text{H}_2\text{SO}_4$

Strong Bases $\text{NaOH}, \text{KOH}, \text{Ca}(\text{OH})_2$ (usually Group I and II soluble hydroxides)

Unless told otherwise you can assume other acids and bases are WEAK.

What can be said about the equilibrium constants for weak electrolytes?

2) DILUTE AND CONCENTRATED

Used to describe how many moles of the acid and base are in one litre of the solution.

A solution with less than 1.2 mol L^{-1} could be called dilute. If it has more than 1.2 mol L^{-1} it could be called concentrated.

Using the words: (acid or base), (weak or strong) and (dilute or concentrated), CLASSIFY the following:

i) $6 \text{ mol L}^{-1} \text{CH}_3\text{COOH}$ is a _____

ii) $0.1 \text{ mol L}^{-1} \text{HNO}_3$ is a _____

iii) $12 \text{ mol L}^{-1} \text{KOH}$ is a _____

iv) $1 \text{ mol L}^{-1} \text{NH}_3$ is a _____

3) DIPROTIC

The H^+ ion can be described as a proton. Why?

So a diprotic acid can give _____ protons when it dissolves in water.

The general name of an acid that gives 3 H^+ is a _____ acid.

DESCRIBING THE BEHAVIOUR OF ACIDS AND BASES

There are 2 models used to describe the behaviour of acids and bases.

ARRHENIUS

This is the simplified view of acids and bases used in the definitions on Page 1. That is, acids make hydrogen ions in water and bases make hydroxide ions in water. Restricted to reactions in solution. Many salts are not neutral and Arrhenius can be used to show how some ions when dissolved in water will produce H^+ or OH^- ions.

Hydrolysis is a reaction with water. Anions (_____) derived from strong acids are neutral and cations (_____) derived from strong bases are neutral.

These ions do not hydrolyse and remain as neutral ions in water. Most other ions will react with water to produce either an acid or a base.

eg F^- is a base. Write an equation to show this.

eg HSO_4^- is an acid. Write an equation to show this.

eg NH_4^+ is an acid. Write an equation to show this.

eg CH_3COO^- is a base. Write an equation to show this.

eg Al^{3+} is an acid. Write an equation to show this.

	NEUTRAL	ACIDIC	BASIC
ANION	From strong monoprotic acids.		
CATION	From strong bases.		

Classify the following chemicals as acid, base or neutral when dissolved in water. Are they strong electrolytes? Are they strong or weak acids and bases?

Salt	Strong Electrolyte?	Acid or Base or Neutral?	Strong or weak acid/base?
sodium chloride			
aluminium chloride			
calcium sulfide			
lithium acetate			
barium hydrogen sulfate			
iron (III) nitrate			
potassium carbonate			
ammonium bromide			

i) Ammonium acetate is a base, explain why. _____

ii) All the non-neutral chemicals above are either weak acids or bases. Using nitric acid and ammonium nitrate as examples, write equations to illustrate the difference between a strong acid and a weak acid. Explain the difference in strength in terms of the conductivity of the resulting solution.

BRONSTED-LOWRY

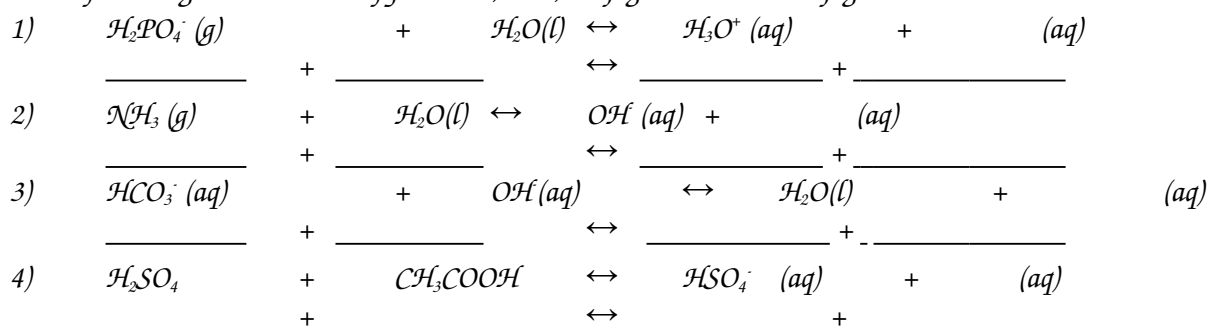
A broader view of acids and bases. Any reaction in which hydrogen ions are transferred from one species to another is an acid-base reaction. The acid _____ the hydrogen ion and the base _____ the hydrogen ion.

Now as products there is the potential to reverse the reaction and are now opposite in their acid/base nature.

An acid gives up its H^+ to become the conjugate base.

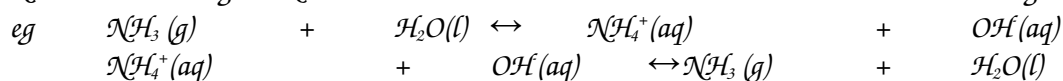
A base accepts the H^+ to become the _____ .

In the following reactions identify the acid, base, conjugate acid and conjugate base.



In the above reactions a reagent may be an acid in one reaction and a base in another. Name the chemical that is acting as an acid in one reactions and a base in another reaction. _____ It could be called a _____ .

Reactions involving weak acids and bases are reversible and both reactions will be occurring at the same time.



CLASSIFYING OXIDES

Oxides can be put into 4 groups. You only need to know about the first three.

- | | | | |
|----|-------------------|---|---|
| 1) | ACIDIC OXIDES | - | React with water to produce acids. |
| | | - | Oxides of most non-metals. |
| 2) | BASIC OXIDES | - | React with acids to make a salt and water. |
| | | - | Oxides of most metals. |
| 3) | NEUTRAL OXIDES | - | Do not react with water to form acids. |
| | | - | Do not react with acids. |
| | | - | Oxides of a few non-metals, H_2O , CO and NO . |
| 4) | AMPHOTERIC OXIDES | - | React with acids and bases. |
| | | - | Oxides of a few metals, <u>Al, Pb, Zn and Cr.</u> ** |

Non-metal oxides are _____ and _____.

Metal oxides are _____ and _____.

Classify the following as acidic, basic, neutral or amphoteric oxides.

i)	<u>Na_2O</u>	ii)	<u>CO_2</u>	iii)	<u>Al_2O_3</u>	iv)	<u>NO</u>
v)	<u>NO_2</u>	vi)	<u>ZnO</u>	vii)	<u>CO</u>	viii)	<u>Fe_2O_3</u>

AMPHOTERIC CHEMICALS

Metals such as zinc, aluminium and chromium are amphoteric (react with both acid and base). Similarly the oxides and hydroxides of these metals are also amphoteric.

BUFFERS

A buffer solution is one whose pH remains almost unchanged if a little acid or alkali is added to it. Buffer solutions are usually one of two types:

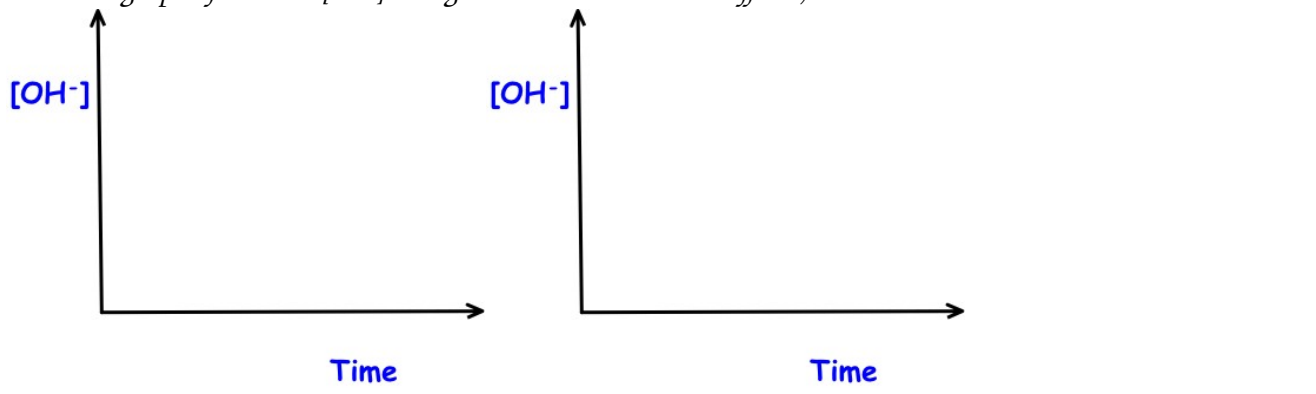
- Acid buffers: made by mixing a weak acid with its conjugate base. Eg CH_3COOH with NaCH_3COO .
- Base buffers: made by mixing a weak base with its conjugate acid. Eg NH_3 with NH_4Cl .

What other mixtures could be used as buffers? (HCN is a weak acid and NaH_2PO_4 is a weak base)

Explain how adding a small amount of acid to buffer i) could produce almost no change in pH. (Use equilibrium)

Explain how adding a small amount of base to buffer ii) could produce almost no change in pH.

Draw the graph of how the $[\text{OH}^-]$ changes in water and in the buffer ii) above.



pH

pH is a measure of how acidic or alkaline a solution in water is. It is a number that comes from the concentration of H^+ ions in the water.

What is the pH of acids? _____. What is the pH of bases? _____. What pH is neutral? _____

Use this formula to
change a hydrogen ion
concentration into pH.

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

What is the pH in the following solutions?

i) $1 \text{ mol L}^{-1} \text{ HCl}$

ii) $7 \text{ mol L}^{-1} \text{ H}_2\text{SO}_4^*$

iii) $10 \text{ mol L}^{-1} \text{ HNO}_3$

Use this formula to
change a pH into a
hydrogen ion concentration

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

What is the $[H^+]$ in the following solutions?

i) HCl ($pH = 4$)

ii) Na_2SO_4 ($pH = 8.5$)

iii) NH_4Cl ($pH = 4.7$)

IONISATION OF MULTIPROTIC ACIDS

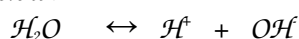
Sulfuric acid is a _____, _____ acid. The ionisation of this acid occurs in 2 steps.

The first ionisation is almost complete but the second ionisation is only partial. What chemical species would be present in a dilute solution of sulfuric acid?

Write the ionisation reactions for phosphoric acid and list which chemical species will be present in a dilute solution of this acid.

IONISATION OF WATER

Water is a weak electrolyte and very few molecules ionise to form H^+ and OH^- ions. The equation for the ionising of water is given below.



It can also be written as a self-ionisation



This may be used to calculate the pH of any solution if either the $[H^+]$ or the $[OH^-]$ is known.

eg 1) In $1 \text{ mol L}^{-1} \text{ NaOH}$:

a) What is the $[Na^+]$?

- b) What is the $[\text{OH}^-]$?
- c) Using the K_w , what is the $[\text{H}^+]$?
- d) Finally what is the pH ?

eg 2) What is the pH in a $0.01 \text{ mol L}^{-1} \text{ Ca}(\text{OH})_2$ solution?

eg 3) What is the pH in pure water (show your working)?

1. If 4.90 g of sulfuric acid is added to water so that the final volume is 100 mL, what will be the pH of the solution, assuming that the sulfuric acid is completely ionised and dissociated?
(3.37 $\times 10^{-4}$)
2. If 25.0 mL of 0.200 M sodium hydroxide solution is added to 30.0 mL of 0.175 M sulfuric acid solution, what is the pH of the mixture? (1.00)
3. 200 mL of 0.0500 M barium hydroxide solution is mixed with 400 mL of 0.200 M nitric acid solution. The mixture is then diluted with water so that the final volume is 6.00 L. What is the pH of the final solution? (2.00)

TITRATION

The word titration refers to a commonly used method of finding the concentration of a _____ solution by comparing it with a known solution. An acid-base titration is good to consider when learning the method, but there are more uses for the technique.

A measured amount from a pipette (called the **aliquot**) of the unknown material in a flask with indicator is usually combined with the known material from a burette. The burette is marked with the volume of solution by a scale with zero on top and 50 mL on the bottom. The burette has a valve at the bottom that can dispense the contained solution.

It is not necessary to start the titration with the known solution level in the burette at the zero mark, but the level must be within the portion of the burette that is marked. The burette is read by getting at eye level to the bottom of the meniscus (curve in the liquid) and comparing the bottom of the meniscus to the marks on the glass. A reading of the burette is taken before and at the end of the titration. The amount of solution used is the difference of the beginning and ending burette reading. This is called the **titre**.

As the titration allows you go from a known concentration to calculating the concentration of an unknown solution you must start with known solution. A **standard solution** is one whose exact concentration is known either by:

- a) Direct weighing out of a substance which is a **primary standard**. The characteristics of a primary standard are:
 - 1) Pure.
 - 2) Does not absorb (hygroscopic) or lose water (efflorescence).
 - 3) Does not react in any way with the air.
 - 4) High molar mass.
- e.g. sodium carbonate (acid-base),
sodium oxalate or oxalic acid (redox).

A solution of a primary standard is prepared in the following manner:

- * The required mass of primary standard is calculated and a sample of approximately this mass accurately weighed.
- * The sample is transferred carefully to a volumetric flask, not losing any solid.
- * The solid is dissolved in distilled water and the solution then made carefully up to the mark.
- * The solution is then thoroughly mixed. Why?

- b) Titration against a primary standard or a previously standardised solution.

In titrations, it is essential to know when the reaction is complete. This point is termed the **equivalence point** (when stoichiometric amounts have reacted, there is no

_____ reactant) and should be signalled by some visible change such as a colour change. In some cases, such as titrations involving the permanganate ion (redox) this change is provided by the reactant itself, but in most cases it is necessary to use an extra reagent to indicate the equivalence point.

Such reagents are called indicators and must be carefully chosen so that the visible change, the **end point** of the titration coincides as nearly as possible with the equivalence point of the reaction. The endpoint is when the indicator changes colour.

Acid-base indicators are substances, which change colour according to the hydrogen ion concentration of the solution to which they are added. The indicators are themselves weak acids. One, or both, of the acid and its conjugate base are coloured.

Litmus	red in acid	No change neutral	blue in base
Phenolphthalein	colourless (below pH 9)		pink (above pH 9).

Methyl Orange	red (below pH 5)	Orange pH 5	yellow (above pH 5)
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The equivalence point may not occur at pH 7, neutral pH, so the appropriate pH indicator must be chosen for the type of acid and base being titrated.

The aim of the titration is to get the indicator end point as close as possible to the reaction equivalence point. This then requires that the indicator chosen is suitable for the reaction is the indicator endpoint is appropriate for the reaction equivalence point.

A **salt** is the combination of an anion (- ion) and a cation (+ ion). Another way to think of a salt is the combination of the anion of a certain acid combined with the cation of a certain base. The neutralization of potassium hydroxide with hydrochloric acid produces _____ and the salt, _____. In a solid salt, the ions are held together by the difference in charge. If a salt dissolves in water solution, it usually dissociates into the anions and cations that make up the salt.

Salts made of the anion of a strong acid and the cation of a strong base will be neutral salts, that is, the water solution with this salt will have a pH of seven.
(example - _____)

Salts made of the anion of a strong acid and the cation of a **weak** base will be acid salts, that is, the water solution with this salt will have a pH of less than seven.
(example - _____)

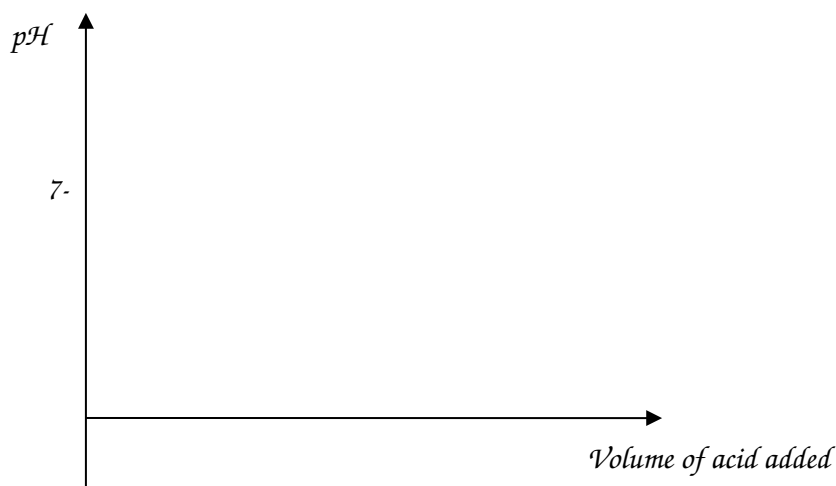
Salts made of the anion of a **weak** acid and a strong base will be an alkali salt. The pH of the solution will be over seven.
(example - _____)

It is more difficult to tell the pH of a salt solution if the salt is made of the anion of a weak acid and the cation of a weak base. The main determining factor is which is stronger the weak acid or the weak base, you are not expected to know so there is no weak acid/weak base titrations.

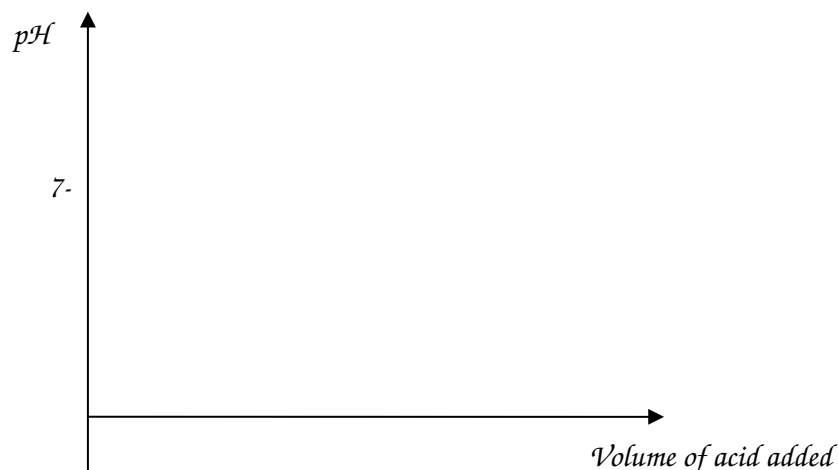
TITRATION CURVES

If the volume of base added to an acid is plotted against pH then the curve can help to decide which indicator should be used.

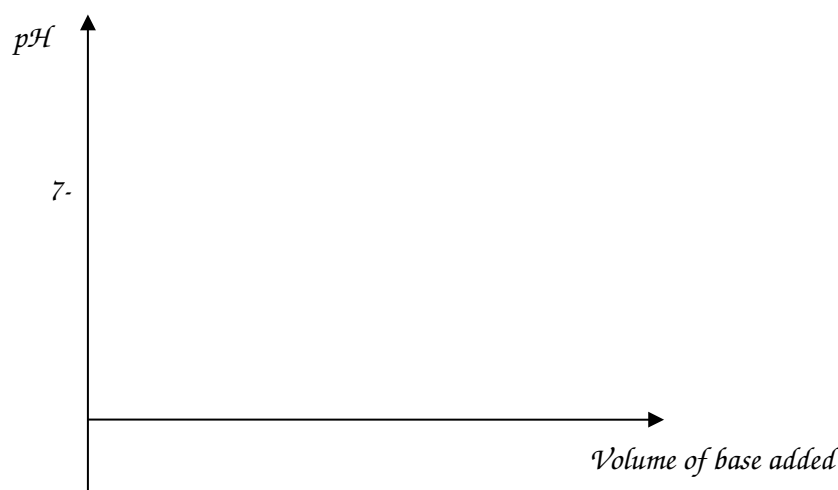
a) Strong acid and strong base



b) Strong acid and weak base



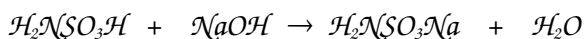
c) Weak acid and strong base



ACID-BASE TITRATION CALCULATIONS

1. 23.45 mL of 0.275 M sodium hydroxide was titrated against 20.0 mL of acetic acid. What was the concentration of acetic acid? (0.322 molL⁻¹)
2. 17.05 mL of 0.247 M barium hydroxide was titrated against 10 mL of nitric acid. What was the concentration of nitric acid? (0.842 molL⁻¹)
3. 35.79 mL of 0.275 M sodium hydroxide was titrated against 15 mL of sulfuric acid. What was the concentration of sulfuric acid in gL⁻¹? (32.2 gL⁻¹)
4. The active ingredient of a drain cleaning powder is sodium hydroxide. A sample of this powder weighing 11.4 g is dissolved in water and made up to 250.0 mL in a volumetric flask. A 20.0 mL sample of the resulting solution required 18.75 mL of 0.320 mol L⁻¹ sulfuric acid for complete neutralisation. You may assume that none of the other ingredients of the drain cleaning powder react with the sulfuric acid.
 - a) How many moles of sulfuric acid were used? (6.00 × 10⁻³ mol)
 - b) What is the concentration of the sodium hydroxide solution? (0.600 molL⁻¹)
 - c) How many grams of sodium hydroxide are contained in 1.00 kg of the drain cleaning powder? (526 g)

5. Sulfamic acid is a strong acid. It is a good primary standard substance for standardising alkalis. The reaction of sulfamic acid with sodium hydroxide is as follows:



2.50 g of sulfamic acid were dissolved in distilled water and diluted in a 250.0 mL volumetric flask. It was found that 20.5 mL of sulfamic acid solution neutralized 20.0 mL of sodium hydroxide solution. Calculate the molarity of the sodium hydroxide solution.
(0.106 mol L⁻¹)

6. 5.10 g of potassium hydrogen phthalate were dissolved in distilled water and diluted to 250.0 mL in a volumetric flask. 20.0 mL of this solution were titrated with 0.105 mol L⁻¹ sodium hydroxide solution, using phenolphthalein as indicator. The volume of sodium hydroxide required to reach end point was 19.05 mL.

Calculate the formula mass of potassium hydrogen phthalate, given that 1 mole of the substance produced 1 mole of H⁺ ions.
(204)

7. 1.10 g of a carbonate of formula, MCO₃, was dissolved in 25.0 mL of 1.0 mol L⁻¹ hydrochloric acid. The resultant solution was titrated with 0.10 mol L⁻¹ sodium hydroxide to determine the excess acid. A volume of 30.0 mL of the hydroxide was needed for neutralisation.

- | | | | | |
|----|--------------------|-----|--|-------|
| a) | Calculate | i) | The formula mass of MCO ₃ . | (100) |
| | | ii) | The relative atomic mass of M. | (40) |
| b) | What element is M? | | | (Ca) |

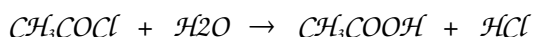
8. 1.50 g of a sample of sodium carbonate decahydrate was dissolved in 25.0 mL (an excess) of 0.50 mol L⁻¹ hydrochloric acid. The resultant solution was titrated with 0.10 mol L⁻¹ sodium hydroxide solution. 25.0 mL of the hydroxide solution were needed for neutralisation. Calculate the percentage purity of the hydrated salt.
(95.3%)

9. 6.00 mL of a laboratory sample of glacial acetic acid were transferred to a 500.0 mL volumetric flask and were then made up to volume with distilled water. 20.0 mL of the diluted solution were titrated with 0.20 mol L⁻¹ potassium hydroxide solution using phenolphthalein as indicator. It was found that 20.9 mL of the hydroxide were needed for neutralization. Calculate:

- | | | |
|----|--|-----------------------------|
| a) | The mass of acetic acid in 1 litre of glacial acetic acid. | (1045 g L ⁻¹) |
| b) | The molarity of glacial acetic acid. | (17.4 mol L ⁻¹) |

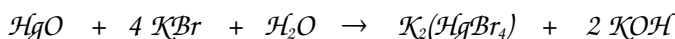
10. 3.00 g of sodium oxide, Na₂O, and 2.00 g of potassium oxide, K₂O, were dissolved in exactly 250.0 mL of distilled water. 25.0 mL aliquots of the solution were titrated with 0.50 mol L⁻¹ hydrochloric acid using methyl red as indicator. Calculate the volume of hydrochloric acid required for neutralisation.
(27.8 mL)

11. Acetyl chloride (also known as ethanoyl chloride) is completely hydrolysed in water to form ethanoic acid and hydrochloric acid:



A certain mass of acetyl chloride was dissolved in distilled water and then diluted to 250.0 mL in a volumetric flask. 20.0 mL of this solution required 21.50 mL of 0.105 mol L⁻¹ sodium hydroxide solution for neutralisation. Calculate the mass of acetyl chloride dissolved.
(1.11 g)

12. Mercury (II) oxide reacts quantitatively with potassium bromide according to the equation:



The potassium hydroxide generated by the reaction can be titrated with a solution of hydrochloric acid. A 2.00 g sample of impure mercury (II) oxide was dissolved in a solution of potassium bromide and was then diluted to 250.0 mL in a volumetric flask. 20.0 mL of the solution, when titrated with 0.100 mol L⁻¹ hydrochloric acid, required 14.0 mL of hydrochloric acid for neutralisation. Calculate the percentage purity of mercury (II) oxide. (Assume that the impurity had no reaction with HCl or NaOH). (94.8%)

13. When dissolved in water, one mole of a hydrated salt, B.xH₂O, produces two moles of hydroxide ions. 5.35g of the hydrated salt were dissolved in water and then diluted to 250.0 mL in a volumetric flask. Using methyl orange as indicator,

20.0 mL aliquots of the solution required an average of 22.0 mL of 0.102 mol L⁻¹ hydrochloric acid for neutralization.

- Calculate the molarity of the solution. (5.61 × 10⁻² mol L⁻¹)
- What is the formula mass of the hydrated salt. (382 g)
- If the anhydrous salt has a formula mass of 202, calculate the value of x. (10)

14. A student analysed a brand of vinegar bought from a supermarket in the following way: 25.0 mL of the vinegar were diluted to 250.0 mL in a volumetric flask. Using phenolphthalein as the indicator, 20.0 mL of the solution required 11.50 mL of

0.105 mol L⁻¹ sodium hydroxide solution for neutralisation. The density of the vinegar solution was 1.05 g mL⁻¹.

- Calculate the number of moles of acetic acid in 20.0 mL of the solution. (1.21 × 10⁻³ mol L⁻¹)
- What is the molarity of the diluted vinegar solution? (6.04 × 10⁻² mol L⁻¹)
- Calculate the mass of acetic acid in 250 mL of diluted solution. (0.906 g)
- What is the percentage by mass of acetic acid in the original vinegar solution? (3.45 %)

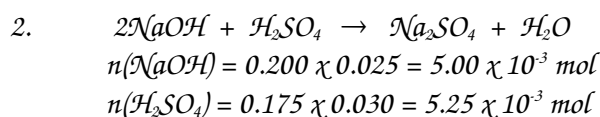
REVISION QUESTIONS

- What mass of HCl must be dissolved in 300 mL of water to give a solution with a pH of 2?
 - What will be the pH of a solution containing 0.0730 g of HCl dissolved in 2.00 L of water?
- What mass of sodium hydroxide must be dissolved in 600 mL of water to give a pH of 13?
 - If 0.600 g of sodium hydroxide is dissolved in 1500 mL of water, what will be the pH of the final solution?
- What volume of water must be added to 20.0 mL of 0.100 M hydrochloric acid to give a solution with a pH of 3?
- What volume of HCl gas (at STP) must be added to 1.00 L of water to produce a solution whose pH is 4?
- Classify each of the species below according to whether (under normal aqueous conditions) it is a Bronsted-Lowry base, a Bronsted-Lowry acid or an amphoteric substance. In each case justify your choice.
 - HCl
 - H₂O
 - NH₃
 - H₂S
 - H₃O⁺
 - HCO₃⁻

6. Identify the species acting as Bronsted-Lowry acids in the following reactions:
- $\text{NH}_4^+(\text{aq}) + \text{HS}(\text{aq}) \leftrightarrow \text{NH}_3(\text{aq}) + \text{H}_2\text{S}(\text{aq})$
 - $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq}) \leftrightarrow \text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{H}_2\text{PO}_4^-(\text{aq}) + \text{NH}_3(\text{aq}) \leftrightarrow \text{HPO}_4^{2-}(\text{aq}) + \text{NH}_4^+(\text{aq})$
 - $\text{H}_2\text{O}(\text{l}) + \text{HF}(\text{aq}) \leftrightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{F}^-(\text{aq})$
7. For each of the aqueous solutions listed below, determine $[\text{H}_3\text{O}^+]$; $[\text{OH}^-]$ and pH.
- 0.0010 M HCl
 - 0.0050 M Ba(OH)₂
 - 300 mL of an aqueous solution of barium hydroxide containing 2.57 g of dissolved Ba(OH)₂.
8. Ammonium carbonate dissolves in water according to the equation:
- $$(\text{NH}_4)_2\text{CO}_3(\text{s}) \rightarrow 2 \text{NH}_4^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \quad \text{.....(1)}$$
- Following dissolution, two further reactions occur:
- $$\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \quad \text{.....(2)}$$
- $$\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq}) \quad \text{.....(3)}$$
- In what way do reactions (2) and (3) differ from reaction (1)?
 - Indicators show that an aqueous solution of (NH₄)₂CO₃ is basic, and pH tests confirm this. Account for this observation in view of the fact that there are twice as many NH₄⁺ ions (which may produce H₃O⁺) as there are CO₃²⁻ ions (which may produce OH⁻) in any sample of ammonium carbonate.
9. Explain, using equations, why aqueous solutions of sodium carbonate, sodium sulfide and sodium ethanoate (sodium acetate) all have pH values greater than 7.
10. Find the pH of:
- An aqueous solution containing 0.0730 g of HCl in a total volume of 2.00 L.
 - An aqueous solution containing 0.800 g of NaOH in a total volume of 2.00 L.
11. A 10.0 mL sample of 0.00500 M Ca(OH)₂ is diluted to 100 mL with water.
- What is the pH of the diluted solution?
 - What change occurs in the pH of the solution due to the dilution?
 - What mass of Ca(OH)₂ is present in the dilute solution?
 - What volume of CO₂, at 25°C and 1.10 × 10⁵ Pa must be bubbled through the diluted solution in order to convert the OH⁻ ions into CO₃²⁻ ions?

ANSWERS TO PAGE 10

1. $n(\text{H}_2\text{SO}_4) = 4.9/98.0616 = 0.0499$
 $n(\text{H}^+) = 2 \times n(\text{H}_2\text{SO}_4)$ (assuming 100% ionisation of H⁺ both ions).
 $= 9.99 \times 10^{-2} \text{ mol}$
 $c = n/V$
 $= 0.0998/0.1$
 $[\text{H}^+] = 0.999$
- $\text{pH} = -\log[\text{H}^+]$
 $= -\log[0.999]$
 $= 3.37 \times 10^{-4}$

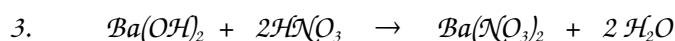


For all of the NaOH to react require $2.50 \times 10^{-3} \text{ mol}$ of acid, hence sulfuric acid is in excess.

$$n(\text{H}_2\text{SO}_4) \text{ in excess} = 5.25 \times 10^{-3} - 2.50 \times 10^{-3} = 2.75 \times 10^{-3}$$

$$\begin{aligned} n(\text{H}^+) \text{ excess} &= 2 \times n(\text{H}_2\text{SO}_4) \quad (\text{assuming } 100\% \text{ ionisation of } \text{H}^+ \text{ both ions}). \\ &= 2 \times 2.75 \times 10^{-3} \\ &= 5.5 \times 10^{-3} \end{aligned}$$

$$\begin{aligned} \text{pH} &= -\log[\text{H}^+] \quad [\text{H}^+] = 5.5 \times 10^{-3} / 0.055 \quad (\text{total } V = 30 + 25 \text{ mL} = 0.055 \text{ L}) \\ &= -\log[0.100] \quad = 0.1 \text{ M} \\ &= 1.00 \end{aligned}$$



$$n(\text{Ba}(\text{OH})_2) = 0.0500 \times 0.200 = 0.0100 \text{ mol}$$

$$n(\text{HNO}_3) = 0.200 \times 0.400 = 0.0800 \text{ mol}$$

To neutralise all $\text{Ba}(\text{OH})_2$, need 0.0200 mol of $\text{HNO}_3 \Rightarrow \text{HNO}_3$ is in excess by 0.0600 mol

$$n(\text{H}^+) \text{ excess} = 0.0600 \quad [\text{H}^+] = 0.06/6 = 0.01 \text{ M}$$

$$\text{pH} = -\log[0.0100] = 2.00$$

ANSWERS TO REVISION QUESTIONS

1. a) $\text{pH} = 2, [\text{H}^+] = 10^{-2} \text{ M}$
 $n = cV = 10^{-2} \times 0.300 = 3.00 \times 10^{-3}$
 $m(\text{HCl}) = n \times M = 3.00 \times 10^{-3} \times 36.45 = 0.110 \text{ g}$
 b) $n(\text{HCl}) = m/M = 0.73/36.45 = 2 \times 10^{-3} \text{ mol}$
 $[\text{H}^+] = n/V = 2 \times 10^{-3} / 2 = 1.00 \times 10^{-3}$
 $\text{pH} = -\log[1.00 \times 10^{-3}] = 3.00$

2. a) For $\text{pH} = 13$, require $[\text{H}^+] = 10^{-13} \text{ M}$
 Since $[\text{H}^+][\text{OH}^-] = 10^{-14}$, then $[\text{OH}^-] = 10^{-1} \text{ M}$
 $n(\text{NaOH}) = cV = 0.1 \times 0.6 = 0.0600$
 $m(\text{NaOH}) = nM = 0.0600 \times 40 = 2.40 \text{ g}$
 b) $n(\text{NaOH}) = 0.6/40 = 0.15 \text{ mol} \quad [\text{OH}^-] = 0.0100 \text{ M}$
 Since $[\text{H}^+][\text{OH}^-] = 10^{-14}$, then $[\text{H}^+] = 10^{-12} \text{ M}$
 $\text{pH} = -\log[10^{-12}] = 12.0$

3. For $\text{pH} = 3$, require $[\text{H}^+] = 10^{-3} \text{ M}$
 Using $c_1V_1 = c_2V_2 \quad V_2 = 2.00 \text{ L}$
 Need to add 1.98 L of water to the 20 mL HCl solution to make 2 L.

4. For $pH = 4$, require $[H^+] = 10^{-4} M$
 Assuming that all the HCl added dissolves: $n(HCl) = 10^{-4}$
 At S.T.P. $n = 10^{-4}$ $V = n \times 22.4$

$$= 10^{-4} \times 22.4$$

$$= 2.24 \times 10^{-3} L \quad (2.24 mL)$$
5. a) HCl , acid, Hydrogen ions produced in solution / donates protons

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

 b) Water – amphoteric
 Acting as a base: $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
 Acting as an acid: $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
 c) NH_3 -Base, accepts protons $NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
 d) H_2S -acid, donates protons $H_2S(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + HS^-(aq)$
 e) H_3O^+ -acid, donates protons $H_3O^+(aq) + H_2O(l) \leftrightarrow H_2O(l) + H_3O^+(aq)$
 f) HCO_3^- - amphoteric
 Acting as an acid: $HCO_3^-(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + CO_3^{2-}(aq)$
 Acting as a base: $HCO_3^-(aq) + H_2O(l) \leftrightarrow H_2CO_3(aq) + OH^-(aq)$
6. a) $NH_4^+ \quad H_2S$ b) $H_2O \quad H_3O^+$ c) $HCO_3^- \quad H_2O$
 d) $H_2PO_4^- \quad NH_4^+$ e) $HF \quad H_3O^+$
7. a) $0.0010 M HCl$ $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
 $[H_3O^+] = [HCl] = 1 \times 10^{-3} M$
 Since $[H^+][OH^-] = 10^{-14}$, then $[OH^-] = 10^{-11}$
 then $pH = -\log(1 \times 10^{-3}) = 3.00$
 b) $Ba(OH)_2 \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$
 $[OH^-] = 2 \times [Ba(OH)_2] = 2 \times 0.005 = 0.0100 M$
 Since $[H^+][OH^-] = 10^{-14}$, then $[H^+] = 10^{-12}$
 then $pH = -\log(10^{-12}) = 12.0$
 c) $n(Ba(OH)_2) = 0.0150 mol$
 $c = [OH^-] = 2 \times [Ba(OH)_2] = 2 \times 0.05 = 0.100 M$
 Since $[H^+][OH^-] = 10^{-14}$, then $[H^+] = 10^{-13}$
 then $pH = -\log(10^{-13}) = 13.0$
8. a) Reaction 1 is a dissociation reaction. The ionic solid is dissolving into ions in aqueous solution. Reactions 2 and 3 are acid-base reactions. In reaction 2, NH_4^+ and H_3O^+ are acting as acids. In reaction 3, H_2O and HCO_3^- are acting as acids.
 b) The equilibrium for reaction 2 lies well to the left, so not much H_3O^+ produced, whereas equilibrium for reaction 3 lies more to the right so more OH^- is produced.
 Overall an excess of OH^- is produced and hence the solution is basic.
9. When Na_2CO_3 dissolves, carbonate ions are produced,

$$CO_3^{2-}(aq) + H_2O(l) \leftrightarrow HCO_3^-(aq) + OH^-(aq)$$

 Na_2S dissolved in water produces the sulfide ion,

$$S^{2-}(aq) + H_2O(l) \leftrightarrow HS^-(aq) + OH^-(aq)$$

 $NaCH_3COO$ dissolved in water produces CH_3COO^- ions

$$CH_3COO^-(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + OH^-(aq)$$

 In all three cases there is enough hydrolysis of the anions to produce OH^- and hence the $pH > 7$.

10. a) $n(\mathcal{H}Cl) = 0.002$ $c(\mathcal{H}Cl) = 0.002/2 = 0.001$
 $p\mathcal{H} = -\log[\mathcal{H}^+] = -\log[0.001] = 3.00$
- b) $n(\mathcal{N}aOH) = 0.02$ $c(\mathcal{N}aOH) = 0.02/2.00 = 0.0100$
 Since $[\mathcal{H}^+][OH^-] = 10^{-14}$, and $[OH^-] = 0.01$ then $[\mathcal{H}^+] = 10^{-12}$
 then $p\mathcal{H} = -\log(10^{-12}) = 12.0$
11. a) $[OH^-] = 2 \times [Ca(OH)_2] = 2 \times 0.005 = 0.010$
 Dilute by a factor of 10, therefore $[OH^-] = 0.00100$ or use $c_1V_1 = c_2V_2$
 Since $[\mathcal{H}^+][OH^-] = 10^{-14}$, then $[\mathcal{H}^+] = 10^{-11}$
 then $p\mathcal{H} = -\log(10^{-11}) = 11.0$
- b) Dilution factor of 10, and $\log[10] = 1$, therefore $p\mathcal{H}$ has gone from $p\mathcal{H} 12$ to $p\mathcal{H} 11$.
- c) $n(Ca(OH)_2) = c \times V = 0.005 \times 0.010 = 5 \times 10^{-5} \text{ mol}$
 $m(Ca(OH)_2) = 5 \times 10^{-5} \times 74.1 = 0.00371 \text{ g}$
- d) $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$
 $n(CO_2) = n(Ca(OH)_2) = 5 \times 10^{-5} \text{ mol}$
 $V = nRT/P$
 $= 1.126 \times 10^{-3} \text{ L}$
 $= 1.13 \text{ mL}$