

Chem!stry

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Essential Notes: Acids, Bases and Salts

Names and Formulae of Some Common Acids, Bases and Salts:

Acids:

Examples include: hydrochloric acid – HCl, nitric acid – HNO₃, sulfuric acid – H₂SO₄, phosphoric acid – H₃PO₄, ethanoic acid – CH₃COOH

Bases / Alkalis:

Examples include: copper(II) oxide – CuO, sodium hydroxide – NaOH, calcium hydroxide – Ca(OH)₂, ammonia – NH₃

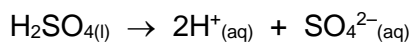
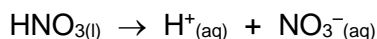
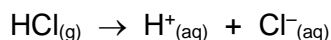
Salts:

Examples include: sodium chloride – NaCl, copper(II) sulphate – CuSO₄, ammonium nitrate – NH₄NO₃

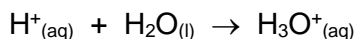
Definitions:

Acid:

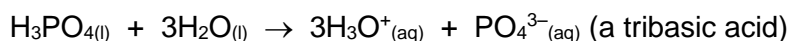
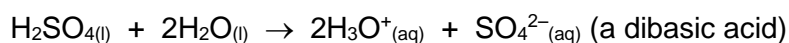
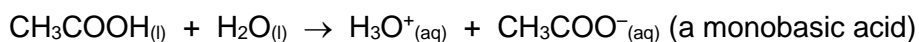
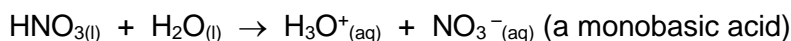
An acid is a chemical that dissolves in water to produce hydrogen ions, H⁺_(aq), as the only positive ion:



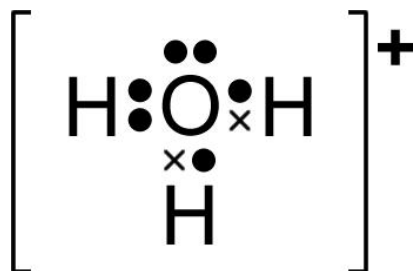
Note: Sulfuric acid is said to be dibasic because 1 mol of the acid produces 2 mol of hydrogen ions. The hydrogen ions that are produced do not exist on their own, but bond to water molecules to form hydroxonium ions, H₃O⁺_(aq):



Therefore, the complete balanced chemical equation for an acid dissolving in water should be written:



The dot and cross diagram of the hydroxonium ion is given below:

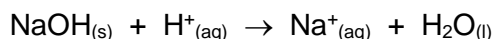
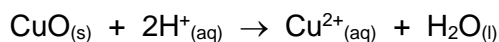


Note: The hydrogen that originated from the acid (shown on the left) is sharing two electrons (two dots) which *both* belong to the oxygen. This special type of covalent bond is known as a dative covalent bond.

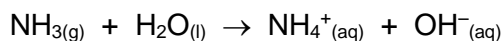
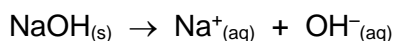
An important fact to arise from this is that an acidic chemical will only exhibit its acidic properties when dissolved in water. Therefore, when writing balanced chemical equations, all acids must be in an aqueous state, (aq).

Base / Alkali:

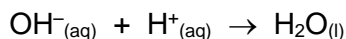
Bases tend to be metal oxides and metal hydroxides. Ammonia, NH_3 , is also basic. As opposed to producing hydrogen ions, which is the property of an acid, a base will accept hydrogen ions to form water as one of the products:



A base which is soluble in water is described as an alkali. Examples of alkalis include sodium hydroxide, NaOH , and ammonia, NH_3 . Alkalis tend to dissolve in water to produce hydroxide ions, $\text{OH}^-_{(aq)}$:



Hydroxide ions (from alkalis) react with hydrogen ions (from acids) to form water. This is known as a *neutralisation reaction*. The ionic equation is:



Salt:

Salts are ionic compounds, usually composed of a positive metal ion bonded to a negative non-metal ion. An exception would be a salt that contains a positive ammonium ion, NH_4^+ , in place of the positive metal ion:

sodium chloride: NaCl composed of Na^+ and Cl^-

copper(II) sulphate: CuSO_4 composed of Cu^{2+} and SO_4^{2-}

ammonium nitrate: NH_4NO_3 composed of NH_4^+ and NO_3^-

Properties of Acids:

- Acids turn blue litmus paper red.
- Acids turn universal indicator red / orange / yellow (depending upon the pH value of the acid).
- Acids have pH values less than 7. A pH of exactly 7 is neutral.

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

Where $[\text{H}^+]$ is the hydrogen ion concentration, mol/dm^3

Example of a pH calculation:

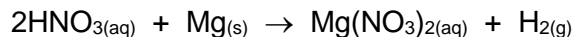
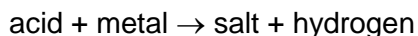
$$\text{If } [\text{H}^+] = 0.001, \text{ then } \text{pH} = -\log_{10} 0.001$$

$$\text{pH} = -(-3)$$

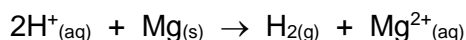
$$\text{pH} = 3.00 \text{ (3 s.f.)}$$

- An aqueous solution of an acid will conduct electricity because it contains mobile ions (charge carrying particles) which are free to move to the electrode of opposite charge. For example, an aqueous solution of sulfuric acid contains a mixture of mobile hydrogen ions, H^+ , which will be attracted to the negative electrode (cathode) and mobile sulphate ions, SO_4^{2-} , which will be attracted to the positive electrode (anode).

- Acids react with metals to produce a salt and hydrogen as the products:



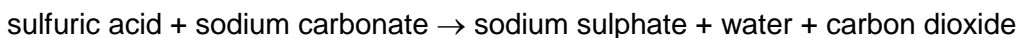
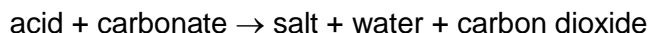
Ionic equation:



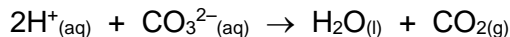
This is a redox reaction because the metal is oxidised while the hydrogen is reduced. This reaction can also be considered as a displacement reaction. The hydrogen of the acid will only be displaced by a metal that is more reactive than hydrogen in the reactivity series, *i.e.* copper, gold and silver will not displace hydrogen from an acid (there will be no observed reaction when copper, gold or silver are added to an acid).

Note: Hydrogen gas extinguishes a burning splint with a “pop” sound.

- Acids react with carbonates to produce a salt, water and carbon dioxide as the products:



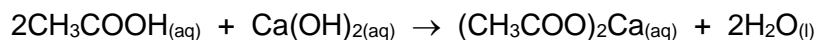
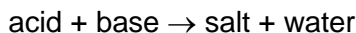
Ionic equation:



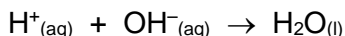
The salt is formed when the hydrogen of the acid is replaced by the metal of the carbonate. This will work for any metal carbonate, including copper(II) carbonate and silver carbonate.

Note: Carbon dioxide gas forms a white precipitate when bubbled through lime water.

- Acids react with bases / alkalis to produce a salt and water as the products:



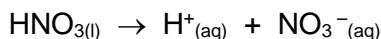
Ionic equation:



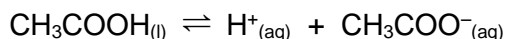
The salt is formed when the hydrogen of the acid is replaced by the metal of the metal oxide / hydroxide. This will work for any metal oxide / hydroxide, including copper(II) oxide and silver hydroxide.

- Acids fall into two categories, strong acids and weak acids.

Strong acids include hydrochloric acid, nitric acid, sulfuric acid and phosphoric acid. These acids fully ionize or dissociate when dissolved in water. For example, if 100 molecules of nitric acid, HNO_3 , are added to water, all of them will ionize to form 100 hydrogen ions, H^+ , and 100 nitrate ions, NO_3^- . There will be no nitric acid molecules left intact:



Weak acids tend to be organic acids (ones that contain carbon) such as ethanoic acid and citric acid. These acids only partially ionize or dissociate when dissolved in water. For example, if 100 molecules of ethanoic acid, CH_3COOH , are added to water, only 5 of them will ionize to form 5 hydrogen, H^+ , ions and 5 ethanoate ions, CH_3COO^- . The remaining 95 ethanoic acid molecules will remain intact:



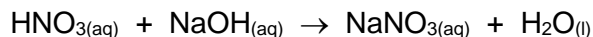
The \rightleftharpoons symbol indicates that the reaction is reversible. This means that once the hydrogen ions and ethanoate ions have been produced, they can combine together again to reform molecules of ethanoic acid.

What is the consequence of this? 1 mole of nitric acid will dissolve in water to form 1 mole of hydrogen ions, but 1 mole of ethanoic acid will dissolve in water to form substantially less than 1 mole of hydrogen ions.

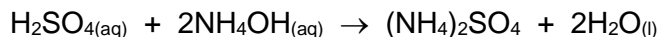
Properties of Bases / Alkalis:

- Alkalis turn red litmus paper blue.
- Alkalis turn universal indicator blue / purple / violet (depending upon the pH value of the alkali).
- Alkalis have pH values greater than 7. A pH of exactly 7 is neutral.
- Alkalis react with acids to form a salt and water as the reaction products:

nitric acid + sodium hydroxide → sodium nitrate + water



sulfuric acid + ammonium hydroxide → ammonium sulphate + water



- Alkalis react with ammonium salts to form a salt, water and ammonia as the reaction products:

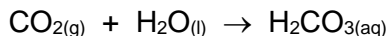
potassium hydroxide + ammonium bromide → potassium bromide + water + ammonia



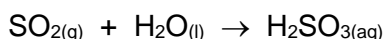
The Nature of Metal and Non-metal Oxides:

Oxides of the chemical elements can be classified in four different ways:

Acidic: Generally, the oxides of the non-metallic elements are acidic. Common examples are carbon dioxide, CO_2 , and sulphur dioxide, SO_2 . Carbon dioxide in the atmosphere dissolves in rain water to form carbonic acid, H_2CO_3 , thus making rain water naturally slightly acidic:

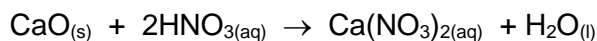


Sulphur dioxide is an atmospheric pollutant produced by the combustion of fossil fuels, such as coal, oil and petrol. It dissolves in rain water to form sulphurous acid, H_2SO_3 , (acid rain) which is harmful to plants, animals and corrodes buildings:



Basic: Generally, the oxides of metallic elements are basic. Common examples include copper(II) oxide, CuO , and calcium oxide, CaO . Both copper(II) oxide and calcium oxide will react with acids to form a salt and water:

calcium oxide + nitric acid → calcium nitrate + water



Amphoteric: These oxides can exhibit both acidic and basic properties, depending upon the chemical environment that they are placed in. For example, if an amphoteric oxide is added to an acidic solution, it will behave as a base. Alternately, if an amphoteric oxide is added to an alkaline solution, it will behave as an acid! Examples of amphoteric oxides include aluminium oxide, Al_2O_3 , lead(II) oxide, PbO , and zinc oxide, ZnO .

Neutral: Neutral oxides have a pH value of 7 when dissolved in pure water. Examples include carbon monoxide, CO, dinitrogen monoxide N₂O, and of course water, H₂O! Note: The neutral oxides all tend to be the *monoxides* of non-metallic elements.

The Properties and Preparation of Salts:

• Solubility:

The table below summarises the solubilities of common salts:

Soluble in Water	Insoluble in Water
All salts of Group I metals, e.g. Na ⁺ and K ⁺ .	All carbonates except those of the Group 1 metals and ammonium carbonate, (NH ₄) ₂ CO ₃ .
All ammonium salts, NH ₄ ⁺ .	All hydroxides and oxides except those of the Group 1 metals, ammonium hydroxide, NH ₄ OH, and barium hydroxide Ba(OH) ₂ .
All nitrates, NO ₃ ⁻ .	All phosphates except those of the Group 1 metals and ammonium phosphate, (NH ₄) ₃ PO ₄ .
All chlorides, Cl ⁻ except silver chloride, AgCl, and lead(II) chloride, PbCl ₂ .	
All sulphates, SO ₄ ²⁻ , except barium sulfate, BaSO ₄ , calcium sulfate, CaSO ₄ , and lead(II) sulfate, PbSO ₄ .	
All hydrogencarbonates, HCO ₃ ⁻ .	

For example:

magnesium chloride, MgCl₂, is soluble, but silver chloride, AgCl, is insoluble.

sodium sulphate, Na₂SO₄, is soluble, but barium sulphate, BaSO₄, is insoluble.

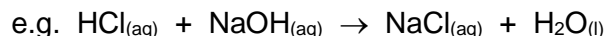
potassium carbonate, K₂CO₃, is soluble, but magnesium carbonate, MgCO₃, is insoluble.

sodium hydroxide, NaOH, is soluble, but copper(II) hydroxide, Cu(OH)₂, is insoluble.

potassium phosphate, K₃PO₄, is soluble, but calcium phosphate, Ca₃(PO₄)₂, is insoluble.

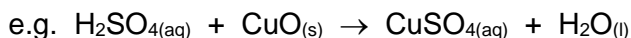
Preparation: There are three possible methods:

1) Reaction of an acidic solution with an alkaline solution to form a soluble salt – titration method



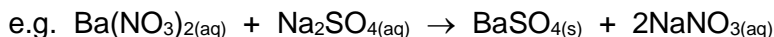
- The acidic solution is pipetted into a flask and the alkaline solution is poured into a burette.
- The acidic and alkaline solutions are then titrated against each other using a suitable indicator.
- When the acid and alkali have exactly neutralised each other, as shown by the colour change of the indicator, the burette reading is taken.
- The titration is then repeated without the indicator, but using exactly the same volumes of acid and alkali that were used in the original titration.
- The resulting solution is heated until it becomes saturated, and the salt then left to crystallise from solution at room temperature.

2) Reaction of acidic solution with an insoluble base or insoluble carbonate to form a soluble salt



- The acidic solution is pipetted into a flask.
- The insoluble base or insoluble carbonate is slowly added to the acid with continuous stirring.
- It is assumed that all of the acid has been neutralised and therefore only the soluble salt (reaction product) is dissolved in solution when no more of the insoluble base or insoluble carbonate is seen to react.
- The excess insoluble base or insoluble carbonate is removed by filtration, leaving a solution of the desired salt.
- The resulting solution is heated until it becomes saturated, and the salt then left to crystallise from solution at room temperature.

3) Reaction of two soluble salts to form an insoluble salt – precipitation method



- A solution of one soluble salt is added to a solution of a second soluble salt to form an insoluble precipitate of a third salt.
- The insoluble precipitate is removed from solution by filtration and rinsed with distilled water to remove any soluble impurities.
- The solid may be pressed between filter papers to aid in its drying.

Note: Never use an insoluble base or insoluble carbonate to form an insoluble salt. You will produce a mixture of two insoluble solids which will be very difficult or even impossible to separate!

Uses of Acids, Bases and Salts:

Sulfuric acid, H_2SO_4 , is a very important acid that is manufactured on an industrial scale in many countries around the world. Sulfuric acid has the following uses:

- Sulfuric acid is used as the electrolyte in car batteries.
- Sulfuric acid is used in the manufacture of detergents that are used to remove grease and oil from clothing and cooking utensils.
- Sulfuric acid is used in the manufacture of fertilisers, such as ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, that add nutrients to the soil for plants to grow.

Sulfur dioxide is an acidic oxide that has the following important uses:

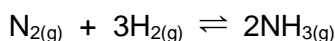
- Sulfur dioxide is used to bleach wood pulp during the manufacture of paper. This is because the sulfur dioxide is a *reducing agent* that will reduce the coloured pigments in the wood pulp, turning them white.
- Sulfur dioxide is used as a preservative in certain foods, such as dried fruits and white wine. It acts as a preservative by killing bacteria and mould that would otherwise grow on the food.

The Industrial Manufacture of Ammonia:

Ammonia, NH_3 , is a very important alkaline gas that is manufactured on an industrial scale in many countries around the world. Ammonia has the following uses:

- Ammonia is used as a raw material in the manufacture of nitric acid, HNO_3 .
- Ammonia is used in the manufacture of fertilisers, such as ammonium nitrate, NH_4NO_3 , that add nutrients to the soil for plants to grow.
- Ammonia is used in the manufacture of explosives, such as 2,4,6-trinitrotoluene (TNT), $\text{C}_7\text{H}_5\text{N}_3\text{O}_6$.
- Ammonia is used in the manufacture of plastics, such as nylon.

Ammonia is manufactured by reacting nitrogen gas and hydrogen gas directly together at a temperature of $450\text{ }^\circ\text{C}$, a pressure of 250 atm. and using an iron catalyst to increase the rate of the reaction:



The reaction is reversible, meaning that nitrogen and hydrogen react together to form ammonia, and ammonia can decompose to form nitrogen and hydrogen.

Note about temperature: A *low* temperature gives a high yield of ammonia, but a low temperature also means that the reaction will be *slow*. A temperature of $450\text{ }^\circ\text{C}$ is a compromise, offering a reasonably good yield of ammonia at an acceptable rate.

Note about pressure: A high pressure gives a high yield of ammonia, but high pressures are both hazardous and expensive to maintain. A pressure of 250 atm. is used as a compromise, offering a reasonably good yield of ammonia at an acceptable cost and level of safety.