**SALTS**

**A salt** *is a substance that is formed when the hydrogen ions in an acid are wholly or partially replaced by positive ion.* Positive ions could be metallic cations or ammonium ion (NH+)

e.g. **H2SO4 (aq) + 2NaOH (aq) Na2SO4 (aq) + 2H2O (l)**

Sodium sulphate (Na2SO4) is the salt formed. Salts are ionic compounds.

Note: Ammonia (NH3) is an unusual base - it does not contain a metal. It forms **ammonium salts**, containing the ammonium ion, NH4+.

**NH3 (aq) + HNO3 (aq)  NH4NO3 (aq)** (ammonium nitrate)

 **Types of salts**

1. **Normal salt**
* Is that which ***doesn’t contain any replaceable, ionisable hydrogen atoms.***

E.g. Sodium chloride(NaCl), potassium sulphate (K2SO4), sodium carbonate (Na2CO3).

These salts are ***neutral in aqueous state***

1. **Acid salts**
* Are salts that ***contain a replaceable ionisable hydrogen atom.***

E.g. sodium hydrogen carbonate (NaHCO3), potassium hydrogen sulphate (KHSO4), sodium di-hydrogen phosphate (NaH2PO4).

They ***have acidic properties in aqueous solution*** due to the presence of replaceable hydrogen ion.

1. **Basic salts**
* Are salts that ***contain hydroxyl ions (-OH)***

E.g. Basic magnesium chloride (Mg(OH)Cl), basic lead (II) carbonate (Pb(OH)2•PbCO3), basic zinc chloride (Zn(OH)Cl), basic copper (II) carbonate (CuCO3•Cu(OH)2)

The presence of the hydroxyl ions (-OH) in these salts that is responsible for the ***basic properties.***

1. **Double salts**
* Are salts in which ***there are two different anions or cations***

E.g. hydrated potassium aluminium sulphate (KAl(SO4)2•12H2O), hydrated ammonium `iron (II) sulphate (Fe(NH4)(SO4).6H2O) and trona (Na2CO3•NaHCO3•2H2O)

**SOLUBILITY OF SALTS IN WATER**

**PROCEDURE**

* Label 4 test tubes 1, 2, 3 and 4. Arrange them in a test tube rack.
* Place about 5cm3 of water into each test tube. Add about a spatulaful of sodium sulphate in test tube 1, sodium chloride in test tube 2, sodium nitrate in test tube 3 and sodium carbonate in test tube 4. Shake the mixture. If the salt appears not to dissolve, warm the mixture. Record your observations.
* Put a tick if the salt is soluble, a cross (X) if the salt is insoluble, **Ss** if the salt is slightly soluble and **Sh** if it is only soluble in hot water.
* Repeat the experiment using similar salts of potassium, ammonium, magnesium, calcium, aluminum, zinc, iron, lead, copper and barium.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Cation/Anion** | **Sulphate** | **Chloride** | **Nitrate**  | **Carbonate**  |
| Sodium  |  |  |  |  |
| Potassium  |  |  |  |  |
| Ammonium  |  |  |  |  |
| Calcium  |  ss |  |  |  |
| Magnesium  |  |  |  |  |
| Aluminium  |  |  |  | Do not exist |
| Zinc  |  |  |  |  |
| Iron  |  |  |  | Do not exist |
| Lead  |  | S in warm H2O |  |  |
| Copper  |  |  |  |  |
| Barium  |  |  |  |  |

**CONCLUSIONS**

1. **Carbonates:**
* All are **insoluble** except potassium, sodium and ammonium carbonate
1. **Chlorides:**
* All are **soluble** except silver and lead (II) chloride. Lead (II) chloride is soluble only in hot water
1. **Nitrates:**
* All are **soluble**
1. **Sulphates:**
* All are **soluble** except barium sulphate, lead (II) sulphate. Calcium sulphate is slightly soluble

**Note:** All potassium, sodium and ammonium salts are soluble in water.

The **preparation** of salt crystals in the laboratory **depends on** whether the salt is **soluble or not**in water.

**SOLUBILITY OF BASES IN WATER**

* Soluble bases are known as **alkali.** They include sodium hydroxide and potassium hydroxide.
* Ammonia gas is also soluble in water to form a basic solution, ammonia solution(ammonium hydroxide)
* Oxides of group II elements are ***slightly soluble in water.***

**CaO + H2O (l) Ca (OH)2 (aq)**

**(Quicklime) (Limewater)**

**MgO(s) + H2O (l) Mg (OH)2 (aq)**

 **(Milk of magnesia)**

* Hydroxides and oxides of the rest of the metals are insoluble in water.

**How To Obtain Crystals Of A Salt From Its Salt Solution I.E. Crystals Of Copper (II) Sulphate From Copper (II) Sulphate Solution**

* Put a solution of copper II sulphate in an evaporating dish and place it on a water bath.
* Evaporate the solution until it is about to form crystals.
* Dip a clean glass rod into the solution and hold it up in the air and examine it to find whether crystals forms on it.
* Continue heating until the crystals form on the glass rod.
* Allow the solution to cool slowly to obtain large crystals.

Discution.

* A saturated solution is a solution in which no more solute can dissolve at a given temperature.
* A hot saturated solution cools to form solid particles called ***crystals.***
* Crystallization is the process by which crystals are obtained from a saturated solution.
* Evaporation to dryness is not done so as to obtain large crystals.
* Crystals of some salts contain water molecules and are said to be hydrated.
* Hydrated salts are salts which contain water of crystallization.
* Water of crystallization is the fixed amount of water incorporated within the crystals of a salt.
* Water of crystallization is important in the formation of crystals for hydrated salts.

**METHODS FOR PREPARING SOLUBLE SALTS**

1. **Action of acids on metal carbonates**

**Acid + carbonate** 🡪 **salt + water + carbon dioxide**

* Useful particularly for making salts of more reactive metals, e.g. calcium, sodium.

 **CuCO3(s) + H2SO4(aq)  CuSO4(aq) + H2O(l) + CO2 (g)**

 **ZnCO3(s) + 2HNO3(aq)  Zn(NO3)2(aq) + H2O(l) + CO2 (g)**

1. **Action of acids on metals**

**Acid + metal** 🡪 **salt + hydrogen**

* Not suitable for making salts of metals above magnesium, or below iron/tin in reactivity. i.e. best for preparing salts of moderately reactive metals e.g.

 **Zn(s) + 2HCl(aq) ZnCl2 (aq) + H2 (g)**

 **Fe(s) + H2SO4 (aq) FeSO4 (aq) + H2 (aq)**

* Metals such as sodium and potassium ***react explosively with acids*** and their salts cannot be prepared by this method
* Lead metal when reacted with dilute sulphuric (VI) acid and hydrochloric acid ***forms insoluble coats of lead (II) sulphate and Lead (II) chloride; which coat rest of reactants stopping the reactions.***
* Metals of copper, silver and mercury do not react with acids and hence cannot be prepared by this method. These metals ***are less reactive than hydrogen and thus cannot replace it from acids to form salts.***

**Exp: Preparation of Zinc sulphate salts**.

* Pour 50cm3 dilute sulphuric (VI) acid into a beaker.
* Add zinc powder a little at a time as you stir with a glass rod. Continue adding zinc powder until it is in excess.
* If the reaction is too slow, ***add a little copper (II) sulphate solution and warm gently, but do not boil***.
* Filter the solution and pour the filtrate into an evaporating dish
* Evaporate the filtrate saturation. Test whether the saturation point is reached by dipping a glass rod into the solution and holding it up in air to cool. If crystals form on the tip of the rod, the solution is ready to form crystals. Heating can be stopped to allow the saturated solution to evaporate slowly.

**DIAGRAM KLB PG 88**

**DISCUSSION**

* Zinc reacts with dilute Sulphuric (VI) acid to produce Zinc Sulphate and hydrogen gas. Effervescence stops when all the sulphuric (VI) acid has completely reacted with Zinc powder.
* ***When copper (II) sulphate solution is added to the reacting mixture, it speeds up the rate of the reaction. It is a catalyst. Warming also increases the rate of the reaction by increasing the temperature of the mixture.***
* ***Excess zinc is used to ensure that all the acid reacts completely.*** Filtration removes the unreacted zinc.

Zinc + sulphuric (VI) acid zinc sulphate + hydrogen gas

**Zn(s) + H2SO4 (aq)  ZnSO4 (aq) + H2 (g)**

* The saturated salt solution is allowed to cool for zinc sulphate crystals to form. Crystals incorporate water as they form. Evaporating rapidly leads to formation of powder (ZnSO4).
* ***Evaporating slowly leads to formation of large crystals*** of Hydrated Zinc (II) sulphate (ZnSO4•7H2O)

**NB**: **A saturated solution** – Is a solution which cannot dissolve anymore solute at a given temperature.

**Water of crystallization** – Is the fixed amount of water incorporated chemically within crystal structures of a salt.

* Evaporation should not be done to dryness so as to ***avoid expulsion of water of crystallization which might otherwise lead to formation of smaller crystals.***
* A salt with water of crystallization is said to be hydrated.
1. **Action of acid on an****insoluble base**

**EXP: Preparation of copper (II) sulphate crystals**

* Pour about 20cm3 of dilute sulphuric (VI) acid into a glass beaker and warm gently; do not boil.
* Add copper (II) oxide to the hot acid a little at a time as you stir until no more oxide can dissolve.
* Filter and collect the filtrate. Transfer the filtrate to the evaporating dish and heat to saturation. Stop heating and allow the saturated solution to cool for to form crystals. Dry the crystals between filter papers.

**DISCUSSION**

* When copper (II) oxide is added to the warm acid, ***a blue solution is formed.*** As copper (II) oxide dissolves, the intensity of the blue color increases.
* ***Excess copper (II) oxide is added to ensure that all the acid has reacted.*** This is a neutralization reaction because an acid is reacting with a base.

 **CuO(s) + H2SO4 (aq) CuSO4 (aq) + H2O(l)**

* Since the reaction is slow, ***warming speeds up the reaction.*** Evaporation of the filtrate is carried over a water bath to ***ensure slow evaporation which leads to formation of large crystals.***
* Other salts that can be prepared by this method include;
1. Lead (II) nitrate

**PbO(s) + 2HNO3(aq) Pb(NO3) 2(aq) + H2O(l)**

1. Magnesium sulphate

**MgO(s) + H2SO4(aq) MgSO4(aq) + H2O(l)**

1. Calcium chloride

**CaO(s) + 2HCl(aq) CaCl2(aq) + H2O(l)**

* An insoluble metal oxide, hydroxide or carbonate, often of a Group 2 metal like calcium, magnesium or a Transition Metal like nickel, copper or zinc are suitable for this method.
* Copper metal won't react with acids, but its oxide and carbonate will. Typical common insoluble bases used for preparing soluble salts include *magnesium oxide, Calcium oxide, calcium hydroxide, zinc oxide and zinc hydroxide.*

**d. Action of acid on alkalis**

**Acid + Alkali(****soluble base)**

* This is useful for making salts of reactive metals, and ammonium salts.
* It is different from other methods as both reactants are in solution.
* This means neutralization must be achieved, by adding exactly the right amount of acid to neutralize the alkali. This can be worked out by **titration.**
* Typical examples include:

 **NaOH (aq) + HCl (aq)  NaCl (aq) + H2O (l)**

 **2KOH (aq) + H2SO4(aq)  K2SO4 (aq) + 2H2O (l)**

 **NaOH (aq) + HNO3 (aq)  NaNO3 (aq) + 2H2O(l)**

**e. Preparation of insoluble salts**

* This involves mixing solutions of two soluble salts that between them contain the ions that make up the insoluble salt. It is made by two methods:
	1. **Making a salt by direct combination of elements**
* Aluminum chloride and iron (III) chloride could be prepared by this method.
* The compounds can be made by ***direct combination of the elements to form anhydrous salts*** e.g. if dry chlorine gas Cl2 is passed over heated iron or aluminum, the chloride is produced. This experiment preparation should be done very carefully in a fume cupboard because chlorine gas is poisonous.
	+ **2Al(s) + 3Cl2(g)  2AlCl3(s)**

The aluminium can burn intensely with violet flame, white fumes of aluminium chloride sublime from the hot reacted aluminum and the white solid forms on the cold surface of the flask.

* + **2Fe(s) + 3Cl2(g) ==> 2FeCl3(s)**

The iron (e.g. as steel wool) glows red and brown fumes of iron(III) chloride stream off, the brown solid collects on the cold flask surface.

**Note (i):** Both these chlorides react exothermically and hydrolyze with water to give the metal hydroxide and fumes of hydrogen chloride and so dry conditions are needed. **Note (ii):** Both these chlorides cannot be made in an anhydrous form from aqueous solution neutralization. This is because the compounds contain 'water of crystallizationand on heating, the hydrated salt hydrolyses and decomposes into water, the oxide or hydroxide and fumes of hydrogen chloride, and maybe some impure anhydrous chloride, basically it is a mess in terms of trying to make pure AlCl3 and FeCl3 in this way.

* 1. **Precipitation/double decomposition**

The starting reagents (salts) must be soluble salts. When mixed, they exchange their ions leading to formation of a precipitate.

**Precipitate**;-is a solid (insoluble salt) formed when aqueous solutions are reacted.

**Exp: Preparation of lead (II) sulphate salt**

* Put 10cm3 of lead (II) nitrate in a beaker.
* To the same beaker, add excess magnesium sulphate solution.
* Stir the solution using a glass rod. Let the solid settle then decant the liquid. Wash the solid with distilled water.
* Filter and dry the solid between filter papers

**DISCUSSSION**

* When lead (II) nitrate and magnesium sulphate solutions are mixed, a **white precipitate** is formed.
* The lead (II) ions (Pb2+) and Sulphate ions (SO42-) react to form insoluble lead (II) sulphate solid, which is insoluble salt. Each salt decomposes and exchanges the ions. This is ***double decomposition.***

 Lead (II) nitrate + Magnesium sulphate Lead (II) sulphate + Magnesium nitrate.

 **Pb(NO3)2 (aq) + MgSO4 (aq)  PbSO4 (s) + Mg(NO3)2 (aq)**

**IONIC EQUATIONS**

* Are equations obtained by writing only those ions that undergo change during a chemical reaction.
* In the above reaction, magnesium ion (Mg2+) and nitrate ions (NO3-) are in aqueous state in the beginning and at the end of the reaction; they remain unchanged throughout the reaction.
* The ions that remain unchanged during a chemical reaction are known as ***spectator ions*** and they are omitted when writing ionic equations.

Only lead ions and sulphate ions reacted.

 **Pb2+(aq) + SO42-(aq) PbSO4(s)**

 **(white precipitate)**

**Steps In Writing Ionic Equations**

1. Write a well-balanced chemical equation. Write the correct state symbols.

**Pb (NO3)2 (aq) + MgSO4 (aq)  PbSO4 (s) + Mg(NO3)2 (aq)**

1. Dissociate the salts that are in aqueous state because they have free ions. Those in ***liquid*** ***gaseous*** and ***solid*** states have no free ions and are left as they are.

**Pb2+(aq) + 2NO3-(aq) + Mg2+(aq) + SO42-(aq) PbSO4 (s) + Mg2+(aq) + 2NO3-(aq)**

1. Those ions in aqueous state appearing on both sides of the equation are cancelled out. The ions remaining constitute the ionic equation. The ones cancelled are the *spectator ions*

**Pb2+(aq) + 2NO3-(aq) + Mg2+(aq) + SO42-(aq) PbSO4 (s) + Mg2+(aq) + 2NO3-(aq)**

The ionic equation thus is:

**Pb2+ (aq) +SO42-(aq) PbSO4 (s)**

Other salts that can be prepared by double decomposition include;

1. **Barium sulphate**

**Ba(NO3)2 (aq) + ZnSO4 (aq)  BaSO4 (s) + Zn(NO3)2 (aq)**

**Ba2+(aq) + 2NO3-(aq) + Zn2+(aq) + SO42-(aq) BaSO4 (s) + Zn2+(aq) + 2NO3-(aq)**

**Ba2+(aq) + SO42-(aq) BaSO4 (s)**

1. **Lead iodide**

**Pb(NO3)2 (aq) + 2KI(aq)  PbI2 (s) + 2KNO3 (aq)**

**Pb2+(aq) + 2NO3-(aq) + 2K+(aq)  + 2I-(aq) PbI2 (s) + 2NO3-(aq) + 2K+(aq)**

**Pb2+(aq) + 2I-(aq) PbI2 (s)**

1. **Copper carbonate**
2. **Silver chloride**

**PROPERTIES OF SALTS**

* 1. **Hygroscopic salts**
* When anhydrous copper (II) sulphate is exposed to the atmosphere, the white solid turns blue and damp.
* This is because it ***absorbs water from the atmosphere but does not dissolve in it*.**
* Such salt is said to be **hygroscopic**, and the process is called **hygroscopy**. MgCl2, KNO3, anhydrous CoCl2and aluminum chloride.

NB: Pure sodium chloride is not hygroscopic but common salt becomes damp because ***it contains traces of magnesium chloride***which makes it behave like hygroscopic salt.

* Concentrated sulphuric (VI) acid is also hygroscopic.
	1. **Deliquescent salts**
* When anhydrous calcium chloride is left in the open it absorbs so much moisture from the atmosphere until it makes a solution.
* ***When a salt absorbs moisture from the atmosphere and dissolves in this water to form a solution*** it is said to be **deliquescent** and the process is called **deliquescence.**
* E.g. FeCl3, ZnCl2, Zn(NO3)2 . KOH and NaOH also exhibit deliquescence.
	1. **Efflorescent salts**
* When crystals of hydrated sodium carbonate (Na2CO3.10H2O) are left in the open a white powder slowly forms on the surface.
* This is because the crystal loses some water of crystallization to form powder, (Na2CO3.H2O).
* ***Salts which lose water of crystallization when exposed to the atmosphere***are said to be **efflorescent** and the process is called **efflorescence.**
* Iron (II) sulphate heptahydrate, (FeSO4.7H2O), sodium sulphate decahydrate (Na2SO4.10H2O) are also efflorescent.

**ACTION OF HEAT ON SALTS.**

1. **Action of heat on carbonates.**
* Pure carbonates of sodium and potassium are not affected by heat.
* They are stable because of the high position of the metals in the reactivity series.
* Lithium carbonate is not thermally stable like the other two.
* Thermal stability increases down group I.
* If hydrated, the salts only lose their water of crystallization when heated.

**Na2CO3∙H2O(s) Heat Na2CO3 (s) +H2O (g)**

**K2CO3∙H2O(s) Heat K2CO3 (s) + H2O (g)**

* Other metal carbonates decompose on heating to give off a colorless gas that forms a white precipitate with lime water; carbon (IV) oxide and their corresponding metal oxides.

**CaCO3(s) Heat CaO(s) + CO2 (g)**

**MgCO3(s) Heat MgO(s) + CO2 (g)**

**ZnCO3(s) Heat ZnO(s) + CO2 (g)**

**PbCO3(s) Heat PbO(s) + CO2 (g)**

**CuCO3(s) Heat CuO(s) + CO2 (g)**

* The position of a metal in the reactivity series determines the ease of decomposition of its carbonate. The higher the reactivity, the harder it is to thermally decompose the metal carbonate
* Ammonium carbonate decomposes to form ammonia gas, carbon (IV) oxide and steam.

**(NH4)2CO3 (s)  Heat 2NH3 (g) + CO2(g) + H2O(g)**

* Ammonia gas turns moist red litmus paper to blue.
* Hydrogen carbonates of metal high in the reactivity series decompose on heating to produce corresponding metal carbonates, carbon (IV) oxide and water.

**2NaHCO3 (s) Heat Na2CO3 (s) + CO2 (g) + H2O (g)**

**2KHCO3 (s) Heat K2CO3 (s) + CO2 (g) + H2O (g)**

* Hydrogen carbonates of calcium and magnesium exist only in solution whereas hydrogen carbonates of aluminum and iron do not exist.
1. **Action of heat on nitrates**
* All nitrates decompose on heating.
1. Nitrates of reactive metals decompose to give white residues of metal nitrites and colorless gas that relights a glowing splint; oxygen gas.

**2NaNO3 (s) Heat 2NaNO2 (s) + O2 (g)**

**2KNO3 (s) Heat 2KNO2 (s) + O2 (g)**

1. Nitrates of metals in the middle of the reactivity series decompose to form metal oxides, nitrogen (IV) oxide and oxygen gas.

**2Ca (NO3)2 (s) Heat 2CaO(s) + 4NO2 (g) + O2 (g)**

 **(White solid)**

**2Mg(NO3)2 (s) Heat 2MgO(s) + 4NO2 (g) + O2 (g)**

 **(White solid)**

**2Pb(NO3)2 (s) Heat 2PbO(s) + 4NO2 (g) + O2 (g)**

 **(Orange-hot**

**Yellow-cold)**

**2Zn(NO3)2 (s) Heat 2ZnO(s) + 4NO2 (g) + O2 (g)**

 **(Yellow-hot**

**white-cold)**

**2Cu(NO3)2 (s) Heat 2CuO(s) + 4NO2 (g) + O2 (g)**

**(Blue crystals) (Black residue)**

1. Nitrates of silver and mercury decompose to give metals, nitrogen (IV) oxide and oxygen gas.

**2HgNO3(s)  Heat 2Hg (l) +2NO2 (g) + O2 (g)**

**2AgNO3(s)  Heat 2Ag(s) + 2NO2 (g) + O2 (g)**

NB/ The ease with which metal nitrates decompose on heating increases down the reactivity series of metals.

Ammonium nitrate decomposes to give steam and nitrogen (I) oxide.

**NH4NO3 (s)  Heat 2H2O (g) + N2O (g)**

Ammonium nitrate explodes when heated and it is not advisable heating it.

1. **Action of heat on sulphate**
* Metal sulphates are much more difficult to decompose and do so only on strong heating.

Sulphates of potassium, sodium, calcium and magnesium are stable and are not affected by heat.

* If hydrated, they lose water of crystallization.

**Na2SO4∙10H2O(s) Heat Na2SO4 (s) + 10H2O (l)**

**MgSO4∙7H2O(s) Heat MgSO4 (s) + 7H2O (l)**

**(Epsom salt)**

**CaSO4∙2H2O(s) Heat CaSO4 (s) + 2H2O (l)**

**(Gypsum)**

* Zinc sulphate and copper sulphate decompose on strong heating to form metal oxides water and sulphur (IV) gas which is colorless.
* The gas turns moist blue litmus paper red then bleaches it.

**ZnSO4∙7H2O** (s) **) Heat ZnO(s) + SO2 (s) + H2O(g)**

**CuSO4∙7H2O** (s) **) Heat CuO(s) + SO2 (s) + H2O(g)**

* Iron (II) sulphate crystals give two gases; sulphur (IV) oxide and sulphur (VI) oxide.

**2FeSO4** (s) **) Heat Fe2O3(s) + SO2 (s) + SO3(g)**

**(Green crystals) (Red-brown solid)**

1. **Action of heat on chlorides**
* Most chlorides are generally very stable. Ammonium chloride undergoes thermal dissociation to produce ammonia gas and hydrogen chloride gas.

**NH4Cl(s)  Heat HCl (g)  + NH3 (g)**

Copper (II) chloride yields copper (I) chloride on heating.

**2CuCl2 (s)  Heat 2CuCl(s) + Cl2 (g)**

* Chlorides are very stable because of the smaller atom of chlorine and its high electronegativity.

**Uses of salts**

* Some are fertilizers e.g. calcium ammonium sulphate, (CAN); diammonium phosphate, (DAP) and ammonium nitrate
* Sodium chloride is used as a food additive while sodium hydrogen carbonate is used in baking powder
* Sodium carbonate is used in softening of hard water, making glass and detergents.
* Potassium nitrate is used in making fireworks and gunpowder
* Calcium sulphate, (plaster of Paris) is used in hospitals on people with fractures or dislocations