



C9I_Detect - Anions(Tests)

Aim : To detect anion (acidic radicals) in given inorganic sample.

Requirements: Test tubes, Test tube stand, Test tube holder, Watch glass, Platinum wire, Glass rod, Plastic beaker (100 ml), Spirit lamp.

Theory: Qualitative analysis is a process of identifying a chemical substance. In inorganic qualitative analysis, the process consists of finding out the basic radical (cation), and acidic radical (anion) present in an inorganic compound.

Qualitative analysis of salt involves the following steps:

1. Preliminary test
2. Dry test for cations / anions
3. Wet test for cations and anions

Detection of Anions:

I. Preliminary test:

	TESTS	OBSERVATIONS	INFERENCE
1	Color	Blue Green Pale yellow White	Cu ²⁺ may be present Cu ²⁺ , Cr ³⁺ , Ni ²⁺ may be present. Cd ²⁺ may present Coloured radicals like Cu ²⁺ , Ni ²⁺ are absent
2	Nature: pinch of salt + water (shake well)	Crystalline (Water soluble) Amorphous (Insoluble in water)	Cl ⁻ , Br ⁻ , I ⁻ , NO ₃ ⁻ , SO ₄ ²⁻ or compounds of K ⁺ or NH ₄ ⁺ may be present Insoluble in CO ₃ ²⁻ , S ²⁻ , SO ₃ ²⁻ may be present
3.	Solubility	i) soluble in water ii) Insoluble in water	i) crystalline salts are present ii) Insoluble CO ₃ ²⁻ , S ²⁻ , SO ₂ ²⁻

II. DRY TEST FOR ANIONS:

I	Tests	Observations	Inference
1	Action of dil. HCl (Salt + dilute HCl)	Colourless gas with brisk effervescence	CO ₃ ²⁻ may be present
2	Action of Conc. H₂SO₄ (Salt + conc. H ₂ SO ₄ + heat)	1) Colourless gas giving dense white fumes with NH ₄ OH 2) Yellowish brown gas turning moist starch paper yellow 3) Violet gas turning starch paper blue 4) Light brown fumes but no effect on moist starch paper	Cl ⁻ may be present Br ⁻ may be present I ⁻ may be present NO ₃ ⁻ may be present
3	Action of MnO₂ and Conc. H₂SO₄ Salt + MnO ₂ + conc. H ₂ SO ₄	1) Colorless or Greenish yellow gas that bleaches moist litmus paper 2) Reddish brown gas turning starch paper yellowish brown 3) Light Brown fumes 4) Violet gas turning starch paper bluish black	Cl ⁻ may be present Br ⁻ may be present NO ₃ ⁻ may be present I ⁻ may be present
4	Cu Foil Test: (Salt + copper filings + conc. H ₂ SO ₄ + heat strongly)	Brown gas evolves and the solution turns bluish green and no action on starch paper.	NO ₃ ⁻ may be present

5	Possibility of SO_4^{2-} If all the tests are negative SO_4^{2-} Radical may be present as there is no dry test for sulphate radical		
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Preparation of Original Solution / Water Extract (O.S. / W.E.)

Dissolve 1 gm. crystalline salt in one test tube (20 ml) of water (stir and boil if necessary) in a beaker. If traces remains insoluble or the solution is turbid, filter the solution. This solution is called Water Extract (W.E) or Original Solution (O.S.). Use this solution for wet test of anions.

III. WET TEST FOR ANIONS:

I	Test	Observations	Inference
1.	W.E.+3-4 ml of dil. HNO_3 , till no effervescence, boil, cool and add AgNO_3	i) Curdy white ppt. completely soluble in NH_4OH . ii) Pale yellow ppt. partly soluble or insoluble in NH_4OH	Cl^- present Br^- I^- present.
2.	Distinction between Cl^- Br^- I^- W.E. + Dil. H_2SO_4 , till acidic+ Chloroform+ Cl_2 water in excess, shake vigorously and observe.	i) Chloroform layer colorless. ii) Chloroform layer yellowish brown. iii) Chloroform layer violet.	Cl^- present. Br^- present. I^- present.
3	C.T. For Cl^- (Chromyl Chloride test) Salt+ solid $\text{K}_2\text{Cr}_2\text{O}_7$ (1:3) + Conc. H_2SO_4 , heat, collect the reddish brown vapours of chromyl chloride in another test tube, add NaOH , Shake add CH_3COOH + $(\text{CH}_3\text{COOH})_2\text{Pb}$	Yellow precipitate	Cl^- is confirmed
4	C.T. for Br^- I^- W.E. + MnO_2 +Conc. H_2SO_4 , heat	i) Reddish brown gas ii) Violet gas	Br^- present. I^- present.
5	Test for SO_4^{2-} W.E + Dil. HNO_3 , till acidic, add $\text{Ba}(\text{NO}_3)_2$ C.T. for SO_4^{2-} W.E. + CH_3COOH + $(\text{CH}_3\text{COOH})_2\text{Pb}$	White Ppt. insoluble in Dil. HNO_3 White Ppt. soluble in Ammonium acetate	SO_4^{2-} present SO_4^{2-} is Confirmed
6	Test for CO_3^{2-} W.E. +Dil. HCl C.T. for CO_3^{2-} W.E. + Phenolphthalein	Colorless gas with brisk effervescence turning lime water milky Pink color decolorizes slowly.	CO_3^{2-} present CO_3^{2-} is confirmed
7	W.E. + Conc. H_2SO_4 (add carefully), cool well under tap water, add freshly prepared FeSO_4 solution by the side of the test tube slowly (donot shake)	Brown ring is formed at the junction of two solutions	NO_3^- present and confirmed.



C9A2_Displacement_Reactions

Aim: To study the displacement reactions:

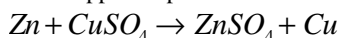
- i) Displacement of copper by zinc
- ii) Displacement of copper by iron
- iii) Displacement of iron by copper
- iv) Displacement of copper by magnesium
- v) Displacement of zinc by copper
- vi) Displacement of zinc by magnesium
- vii) Displacement of magnesium by copper

Requirements: Test tubes, Test tube stand, etc.

Chemicals: $CuSO_4$ Solution, zinc granules, piece of iron wire, piece of copperwire, $FeSO_4$ solution piece of magnesium wire piece of copper wire, $ZnSO_4$ solution $MgSO_4$ Solution.

Theory: - The position of a metal in the reactivity series tells chemists how it will behave in the presence of other metals. Sodium and magnesium can be put at the top of the reactivity series, with copper nearer the bottom. Zinc is higher than copper: if, for instance, zinc is placed in copper sulphate solution, it will dissolve, replacing the copper ions in the solution. Copper metal becomes visible as a deposit on the undissolved zinc. This is an example of a displacement reaction.

Displacement reaction: Displacement Reaction is a chemical process in which one element changes places with another in a compound, that is to say, displaces it. E.g Zinc displaces copper from copper sulphate solution.



Whether or not a metal displaces another in a reaction depends on their relative positions in the reactivity series. A metal that is higher in the series displaces one that is lower. Zinc is above copper and lead, and both of these are above silver. Magnesium is one of the highest metals and can displace almost every other metal, whereas tin is low in the series and displaces few. The reactivity series can help engineers decide which metal is best for a particular job. Since sodium reacts violently with water, they would not consider using it for holding water. Copper, far below sodium in the reactivity series, will not even react with steam, so it can be used for hot-water tanks. Iron, above copper in the series, will react with steam, so it is used, in the form of steel, for cold-water tanks only. An iron hot-water tank would wear out very quickly.

Reactivity series of metals

Metals		Reactions					
		Air or Oxygen	Water	Dilute Acids	Reduction of Oxygen by hydrogen*	Displacement Reactions	Method Of Extraction
Potassium	K	Burn Violently	React with decreasing vigor	React with decreasing vigor	Not reduced	Each metal displaces others lower in the series	Electrolysis
Sodium	Na						
Calcium	Ca	Burn Rapidly	React with steam	Do not react	Readily reduced	Reduction of oxides	Occurs as free elements in rocks
Magnesium	Mg						
Aluminum	Al	Do not react	Do not react	Do not react	Do not react	Do not react	Do not react
Zinc	Zn						
Iron	Fe	Oxidize Slowly	Do not react	Do not react	Do not react	Do not react	Do not react
Lead	Pb						
Copper	Cu	Do not react	Do not react	Do not react	Do not react	Do not react	Do not react
Mercury	Hg						
Silver	Ag	Do not react	Do not react	Do not react	Do not react	Do not react	Do not react
Gold	Au						

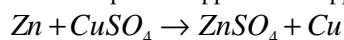
* When heated to no more than $800^{\circ}C$ ($1,440^{\circ}F$)

Procedure:

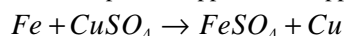
1. Take about 5 ml of copper sulphate solution in a test tube. Add a few zinc granules to the solution and shake the solution for sometime. Observe the color of the zinc granules
2. Take about 5 ml of copper sulphate solution in a test tube. Add a well polished piece of iron wire to the copper sulphate solution. Shake the solution for sometime and observe the color of the iron wire.
3. Take about 5 ml of iron sulphate solution in a test tube. Add a piece of copper wire to the solution and shake the solution for sometime. Observe the color of the copper wire
4. Take about 5 ml of copper sulphate solution in a test tube. Add magnesium wire to the solution and shake the solution for sometime. Observe the color of the magnesium wire.
5. Take about 5 ml of zinc sulphate solution in a test tube. Add a piece of copper wire to the solution and shake the solution for sometime. Observe the color of the copper wire.
6. Take about 5 ml of zinc sulphate solution in a test tube. Add a piece of magnesium wire to the solution and shake the solution for sometime. Observe the color of the magnesium wire.
7. Take about 5 ml of magnesium sulphate solution in a test tube. Add a piece of copper wire to the solution and shake the solution for sometime. Observe the color of the copper wire.

Reactions:

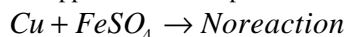
1. Zinc displaces copper from copper sulphate solution.



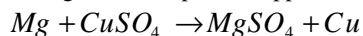
2. Iron displaces copper from copper sulphate solution.



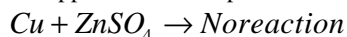
3. Copper does not displace iron from iron sulphate solution.



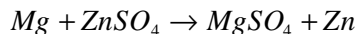
4. Magnesium displaces copper from copper sulphate solution.



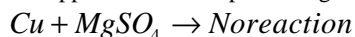
5. Copper does not displace zinc from zinc sulphate solution.



6. Magnesium displaces zinc from zinc sulphate solution.



7. Copper does not displace magnesium from magnesium sulphate solution.



Metals can be arranged in a sequence called the reactivity series, according to the readiness with which they undergo chemical reactions such as burning, or dissolving in acids. A metal generally reacts more vigorously than those that are below it in the series, and it can take their place (or “displace” them) in various compounds or in solution. However, in some reactions, such as reduction reactions, the order of reactivity is reversed

The activity series of cation elements (metals) shows that gold is the least active metal. That should not be surprising, because gold does not tarnish. If we were to consider the Group 1 elements only on the activity list, lithium is the least active and francium is the most active, with each larger element being more active than the smaller one above it on the Periodic Chart. On the other side of the chart we could consider an activity series for anions. Taking just the halogens, the smallest halogen, fluorine is the most active. As the size of the halogen increases down the chart, the activity decreases. If an element is more active than the element of the same sign in an ionic solution, the more active element will replace it.



C9P4_EVOLUTION_AND_ABSORPTION

Aim: To study the evolution & absorption of energy in a chemical reaction.

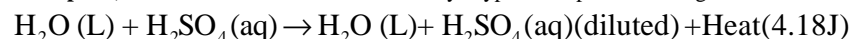
- Dissolution of NaOH in water.
- Dissolution of NH_4Cl in water.
- Dissolution of NaNO_2 in water.

Requirements: Small beakers, thermometer, water, NaOH, NaNO_2 , NH_4Cl .

Theory: When change in heat energy is observed the reactions are known as Thermochemical Reactions. They are of two types

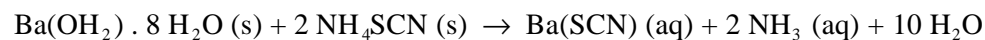
i) Exothermic: Some chemical reactions evolve heat, resulting in a warming of the surroundings (Exothermic Reactions). In exothermic reaction there is significant rise in the temperature in the beaker and the beaker become hot than before. Most of the combination reactions are exothermic in nature as heat is liberated while compounds are formed.

Example: i) When we add **Acid to Water** slowly. Typical temperature change will be from around 22°C to 93°C . The equation for this reaction is:



ii) Endothermic: Some chemical reactions absorb heat from the surroundings, resulting in a cooling of the surroundings (Endothermic Reactions). An endothermic reaction lowers temperature in beaker so that the beaker will freeze and become cooler than before. Breaking of the bonds requires supply of heat hence decomposition reactions are classified into endothermic reaction.

Example: When we add the Ammonium thiocyanate into the flask containing the barium hydroxide. Stir chemicals together by twirling them in the flask. It absorbs heat from surroundings and become cooler. The two solids will gradually react as follows:



Definitions

- Activation Energy of a reaction, E_a , is the minimum amount of energy reactant molecules must possess in order to form products. On an Energy Profile, the activation energy is measured from the energy of the reactants to the peak of the energy profile diagram.
The lower the activation energy, the faster the reaction will proceed.
- Enthalpy change**, ΔH , is the amount of energy absorbed or released by the reaction.
On an Energy Profile, the enthalpy change is measured from the energy of the reactants to the energy of the products.
- Exothermic** reaction, energy is released by the reaction, the energy of the reactants is greater than the energy of the products. (ΔH is negative)
- Endothermic** reaction, energy is absorbed by the reaction, the energy of the reactants is less than the energy of the products. (ΔH is positive)
- Catalysts speed up the rate of reaction by lowering the activation energy without being consumed by the reaction.
A catalyst does not alter the enthalpy change value for the reaction

Procedure:

- Take about 25 ml of water in a beaker. Using the thermometer record the temperature of the water taken (T_1).
- Add a small portion of NaOH & stir the solution. Record the temperature of water (T_2).
- Take about 25 ml of water in a beaker. Record the temperature of water taken (T_1). Add small amount of NH_4Cl & stir the solution. Record the temperature of water (T_3).
- Take about 25 ml of water and record the temperature of water (T_1). Add small amount of NaNO_2 and stir the solution. Again record the temperature of water (T_4).



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C9_Factors affecting rate of chemical reactions

Aim: To study the effect of the size of the particles and the concentration of reactants on the rate of reaction of calcium carbonate (shahabadstones) and hydrochloric acid.

Requirements: - Downward delivery tube, rubber cork, shah bad stone CaCO_3 , dilute, conc. HCl , freshly prepared

Theory: Rate of chemical reaction can be changed as rate depends upon various factors. Factors affecting rate of reactions are

- i) Nature of reactants
- ii) Size of particles
- iii) Concentration
- iv) Temperature &
- v) Catalyst

Smaller the size of the particles of a solid taking part in a reaction, greater is the rate of reaction. This is evident because in the case of the solids, taking part in a reaction, only the particles available on the surface are available for the reaction. Therefore, the rate of reaction depends on the surface area of the reactant. Hence, finer the particles of the solid the more will be the surface area available, the more favorable the condition for the reaction. Hence in a lab finely powdered chemicals are used instead of the lumps or big pieces. More the concentration of the reactants more will be the speed of the reaction i.e. The rate of reaction is proportional to the concentration of the reactants. A good example of the effect of concentration is burning of candle in a gas jar than air. This is because the concentration of oxygen is nearly five times on the gas jar, as compared to the air. Thus, we realize the importance of higher percentage of nitrogen in air which controls the oxidation reactions in nature.

Reactants and products:

A chemical reaction involves one or more reactants and products:

In the example above, zinc in a solid form, as indicated by the symbol (s), reacts with sulphuric acid in an aqueous solution, as indicated by the symbol (aq). These are the **reactants**. The reactants react to form **products** written to the right of the arrow. In the above example, the products are an aqueous solution of the salt zinc sulphate and hydrogen in gaseous form, which is indicated by the symbol (g).

Definition of the reaction rate:

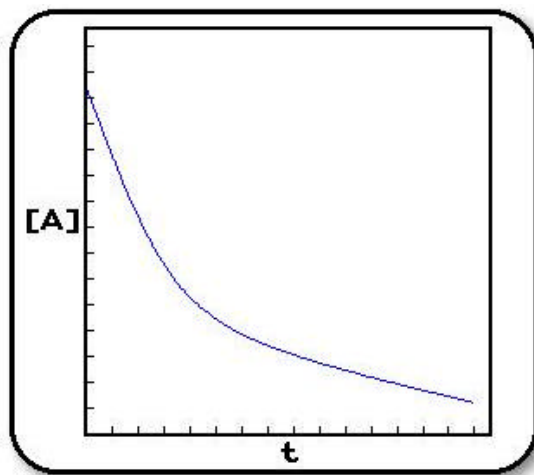
The rate of a chemical reaction is defined as the **change in the concentration of one of the reactants or products in unit time**

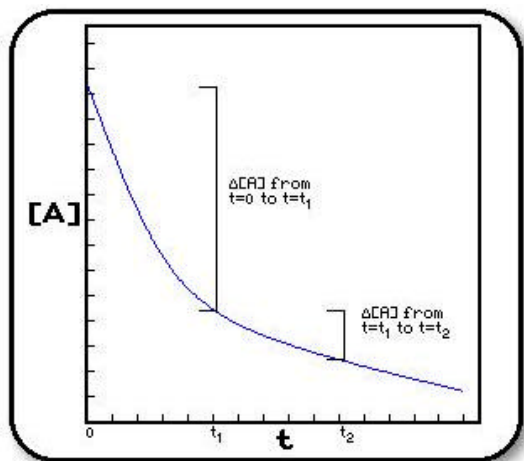
For example, in the reaction above, we could express the rate of the reaction as the change in the concentration of H_2SO_4 ($\Delta[\text{H}_2\text{SO}_4]$) in a certain time interval :

Let's consider a general reaction: $\text{A} \rightarrow \text{B}$:

If one plots the concentration of A, [A] against the time, t, one typically obtains a curve such as the one shown on the right:

The fact that the graph is not a straight line tells us the rate of the reaction changes with time.





In the time $t = 0$ to $t = t_1$, there is a comparatively greater change in $[A]$ than in the same time interval $t = t_1$ to $t = t_2$.

The rate of a reaction is therefore always linked to the specific time interval in which the concentration measurements were made!

Initial rates:

The **initial rate** of a reaction is the rate right at the start of the reaction, when $t = 0$. It is calculated by determining the slope of the tangent to the curve at time $t = 0$:

looking at the curve again, we see that its steepest part occurs at the start of the reaction. That will be when the rate will have its highest value.

$$\begin{aligned} \text{Slope} &= \frac{\Delta[A]}{\Delta t} \\ &= \frac{([A] \text{ at time } = t) - ([A] \text{ at time } t = 0)}{t} \\ &= \frac{[A]_t - [A]_0}{t} \end{aligned}$$

Factors affecting reaction rates:

Reaction rates are greatly influenced by a number of factors:

- **CONCENTRATION:** the greater the concentration of a reactant, the greater the rate of the reaction.
- **PRESSURE:** increases the rates of reactions which take place with the reactants in the gaseous phase.
- **TEMPERATURE:** The higher the temperature, the higher the reaction rate.
- **CATALYSTS** are substances which speed up reaction rates without undergoing a permanent change in the process. The process is called **CATALYSIS**, and it may be **HOMOGENEOUS**, when both the reactants and the catalyst are in the same phase (for example, when both are in solution), or **HETEROGENEOUS**, when the reactants and catalyst are in different phases (for example, gaseous reactants and a solid catalyst.)
- **BOND TYPES:** Reactions between ionic compounds are usually very much faster than those involving compounds where covalent bonds have to be made or broken.

STATE OF SUBDIVISION: This is very important when solids are involved. The more finely divided the solid is, the faster the reaction will take place. In fact, for solids, it is the surface area, and not the concentration, which affects the rate.

Procedure:

1. Take 2g of coarse pieces of shah bad stones in a downward delivery tube.
Add about 5 ml of dilute
2. Fix a cork on the delivery tube. Insert the open end of the delivery tube into the limewater taken in another test tube.
Observe the evolution of the CO_2 and the rate at which the lime water turns milky.
3. Repeat the experiment with conc. HCl and powdered shahbad stones.
Take 2 g of powdered $CaCO_3$ and add about 5ml of dil. HCl to it and observe the evolution of the CO_2 and the rate at which the limewater turns milky.
4. Take 2 g of powdered $CaCO_3$ and add about 5ml of conc. HCl to it and observe the evolution of the CO_2 and the rate at which the limewater turns milky.

