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CHEMISTRY

REVISION NOTES

GROUP - 1 & 2

Group 1 & 2

Atomic radius

As one goes down the group, the atomic radius increases because the atoms have more shells of electrons, making the atom bigger.

Melting points

The melting points of elements decrease as you move down the group. This is because the metallic bonding becomes weaker as the atomic size increases. As the distance between the positive ions and delocalised electrons increases, the electrostatic attractive forces between them weaken.

1st ionisation energy

The first ionisation energy is the energy required to remove the outermost electron from an atom. This energy is higher for atoms with a stronger attraction between the nucleus and the electrons. As the outermost electrons are located in shells that are successively further from the nucleus, they are held more weakly.

Additionally, the outer shell electrons are shielded from the attraction of the nucleus by the repulsive force of inner shell electrons.

Group 1 & 2 reactions

The reactivity of group 1 & 2 metals increases down the group.

Group 1 and 2 metals will react with oxygen to form metal oxides. For instance, when four atoms of sodium (4Na) react with one molecule of oxygen (O_2), they form two molecules of sodium oxide ($2Na_2O$).

The balanced chemical equation for this reaction is

$4Na + O_2 \rightarrow 2Na_2O$.

The metals in group 2 react vigorously with oxygen. Magnesium burns with a bright white flame.

The reaction of 2Mg and O_2 produces 2MgO. MgO is a white solid with a high melting point due to its ionic bonding.

Mg will also react slowly with oxygen without a flame.

Mg ribbon will often have a thin layer of magnesium oxide on it formed by a reaction with oxygen.

$2Mg + O_2 \longrightarrow 2MgO$

This needs to be cleaned off with emery paper before doing reactions with Mg ribbon.

If testing for reaction rates with Mg and acid, an uncleaned Mg ribbon would give a false result because both the Mg and MgO would react but at different rates.

$Mg + 2HCl \longrightarrow MgCl_2 + H_2$
$MgO + 2HCl \longrightarrow MgCl_2 + H_2O$

Exam point – learnt the difference between the reaction of magnesium with steam and warm water.

- Magnesium reacts slowly with cold water to form solution of magnesium hydroxide &hydrogen:

$$Mg(s) + 2H_2O_{(l)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

- The reaction with steam occurs readily, however, & the products are magnesium oxide & hydrogen:

$$Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2(g)}$$

The other Group 2 metals react with cold water with increasing force as we move down the group, forming hydroxides. The general equations is as follows:

$$\boxed{\text{METAL} + \text{WATER} \longrightarrow \text{METAL HYDROXIDE} + \text{HYDROGEN}}$$

$$Ba + 2H_2O \longrightarrow Ba(OH)_2 + H_2$$

Upon adding reactive metals to water, observations are as follows:

- One would observe fizzing.
- The metal dissolves more vigorously down the group.
- The solution heats up.

With calcium, a white precipitate appears, while less precipitate forms a down group with other metals.

Reactions with Chlorine

The group 1 and 2 metals will react with chlorine as follows:

$2\ Na + Cl_2 \rightarrow 2\ NaCl$
$Mg + Cl_2 \rightarrow MgCl_2$

Reactions of the oxides of Group 1 and 2 elements with water

Ionic oxides act as bases because they accept protons from water molecules to form hydroxide ions, making them Bronsted-Lowry bases.

For instance, the reaction between sodium oxide and water produces sodium hydroxide, which is highly alkaline and has a pH of 14. All group 1 oxides react similarly and form highly alkaline solutions.

The balanced equation for the reaction is:

$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$.

The resulting solution has a pH of approximately 14.

Similarly, magnesium oxide reacts with water to produce magnesium hydroxide, which has a pH of 9.

The chemical equation for this reaction is: $MgO + H2O \rightarrow Mg(OH)2$.

As a result, the pH of the solution will be basic, typically greater than 9.

However, since magnesium hydroxide is only slightly soluble in water, fewer free hydroxide ions are produced, resulting in a lower pH.

Group 1 and 2 ionic oxides react with water to produce hydroxides. When magnesium oxide reacts with hydrochloric acid, it forms magnesium chloride and water. Similarly, when magnesium hydroxide reacts with nitric acid, it forms magnesium nitrate and water.

Finally, the reaction between calcium oxide and water produces calcium hydroxide, which has a pH of 12.

Thermal decomposition of Group 1& 2 carbonates

The thermal decomposition of carbonates in Groups 1 and 2 exhibits a distinct trend. As you move down each group, the ease of thermal decomposition decreases.

For example, calcium carbonate ($CaCO_3$) in Group 2 breaks down into calcium oxide and carbon dioxide upon heating:

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

A similar trend is observed in other Group 2 carbonates, such as magnesium carbonate ($MgCO_3$):

$MgCO3(s) \rightarrow MgO(s) + CO2(g)$

Thermal decomposition is the process of using heat to break down a reactant into multiple products. Interestingly, Group 2 carbonates become more thermally stable as you go down the group. This is because the larger cations cause less distortion to the carbonate ion, weakening the C-O bond to a lesser extent and making the decomposition less easy.

Experimental investigations into the ease of decomposition involve heating a known mass of carbonate in a side-arm boiling tube. The evolved gas is then passed through lime water, and the time taken for the first appearance of permanent cloudiness in the lime water serves as an indicator of the decomposition rate. These experiments are conducted using the same amount of carbonate, equivalent volumes of lime water, and a consistent Bunsen flame with the tube placed at a constant height above the flame.

It's worth noting that Group 1 carbonates, except for lithium carbonate, do not readily decompose. The +1 charges on Group 1 cations are generally insufficient to induce significant polarisation of the carbonate ion. Lithium is an exception due to the small size of its ion, allowing for a notable polarising effect.

In summary, the thermal stability of carbonates in Groups 1 and 2 follows a clear pattern, with the size of cations playing a crucial role in influencing the ease of thermal decomposition.

Thermal decomposition of group 1 & Nitrates

The ease of thermal decomposition decreases down the group.

The thermal decomposition of Group 1 and 2 nitrates results in the production of Group 2 oxides, oxygen, and nitrogen dioxide gas. When group 2 nitrates are heated, you will observe the evolution of brown gas, which is NO_2. Additionally, the white nitrate solid will melt and form a colourless solution before resolidifying.

Thermal stability of Nitrates

The reason for the change in thermal stability is similar to that for carbonates. Magnesium nitrate decomposes more easily because the Mg^{2+} ion is smaller and has a greater charge density. This causes greater polarisation of the nitrate anion and weakens the N-O bond.

Group 1 nitrates decompose differently than group 2, except for Lithium nitrate. They decompose into a Nitrate (III) salt and oxygen.

The thermal decomposition of sodium nitrate ($NaNO_3$) yields sodium nitrite ($NaNO_2$) and oxygen gas (O_2).

$$2NaNO_3 \rightarrow 2NaNO_2 + O_2$$

Group 2 nitrates and lithium nitrate decompose in the same way according to the following equation:

$$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$$

Solubility of Hydroxides

Group 2 metals, also known as alkaline earth metals, include beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). The solubility of their hydroxides in water generally increases down the group.:

Beryllium Hydroxide ($Be(OH)_2$):

- Beryllium hydroxide is amphoteric but tends to form basic solutions. It is sparingly soluble in water.

Magnesium Hydroxide ($Mg(OH)_2$) – Medical use :

- Magnesium hydroxide is sparingly soluble in water. However, it has low solubility, and its solid form is often used as a suspension (milk of magnesia) for medicinal purposes to treat excessive acidity in the stomach.

Calcium Hydroxide ($Ca(OH)_2$):

- Calcium hydroxide is sparingly soluble in water. It does dissolve to some extent, and the resulting solution is known as **lime water which can used to test for Carbon dioxide.**

$$Ca(OH)_{2(aq)} + CO_{2(g)} \longrightarrow CaCO_{3(s)} + H_2O_{(l)}$$

Strontium Hydroxide ($Sr(OH)_2$):

- Strontium hydroxide is slightly more soluble than the hydroxides of the preceding elements but is still considered sparingly soluble.

Barium Hydroxide ($Ba(OH)_2$):

- Barium hydroxide is more soluble than the hydroxides of the previous elements in the group. It is still considered sparingly soluble but is more soluble than strontium hydroxide or calcium hydroxide.

When mixed with water, Barium hydroxide dissolves easily, and the solution becomes strongly alkaline due to the presence of hydroxide ions. The chemical equation for the reaction is

$Ba(OH)_2$ (s) + H_2O (l) → Ba^{2+} (aq) + 2OH- (aq).

Solubility of Sulphates:

The solubility of sulfates of Group 2 elements, also known as alkaline earth metals, generally decreases as we move down the group.

Here's a general trend in the solubility of sulfates for Group 2 elements:

- Beryllium sulfate ($BeSO_4$) is not able to dissolve in water.

- Magnesium sulfate ($MgSO_4$) is able to dissolve in water.

- Calcium sulfate ($CaSO_4$) is only able to dissolve sparingly in water.

- Strontium sulfate ($SrSO_4$) is only slightly able to dissolve in water.

- **Barium sulfate ($BaSO_4$) is almost completely insoluble in water.**

Medical use of Barium Sulphate:

$BaSO_4$, also known as barium sulfate, is a chemical compound used in the medical field to help diagnose gastrointestinal problems. It is given to patients as a 'Barium meal' before taking x-ray images of their intestines. The barium in the compound absorbs the X-rays, allowing the gut to show up clearly on the X-ray image. Although barium compounds are generally toxic, it is safe to use in this case because the low solubility of $BaSO_4$ means it is not absorbed into the bloodstream.

Barium reaction with Sulfuric acid

When barium metal is reacted with sulfuric acid, the reaction will happen slowly due to the formation of an insoluble compound called barium sulfate. This compound will be produced on the surface of the metal, preventing further reaction with the acid.

The chemical equation for this reaction is:

$$Ba + H_2SO_4 \rightarrow BaSO_4 + H_2.$$

A similar effect occurs with metals that are higher up the group because their solubility increases. However, this effect is less pronounced.

It's worth noting that this reaction doesn't happen with other acids like hydrochloric or nitric acid because they form soluble group 2 salts.

Flame Test

Flame Test for Different Metals:

1. Lithium (Li):
- *Flame Color:* Crimson red

2. Sodium (Na):
- *Flame Color:* Intense yellow

3. Potassium (K):
- *Flame Color:* Lilac or light purple

4. Calcium (Ca):
- *Flame Color:* Orange-red

5. Barium (Ba):
- *Flame Color:* Apple green

6. Copper (Cu):
- *Flame Color:* Blue-green

7. Strontium (Sr):
- *Flame Color:* Crimson red (similar to lithium but deeper)

8. Rubidium (Rb):
- *Flame Color:* Red-violet

9. Cesium (Cs):
- *Flame Color:* Blue-violet

Note:
- Flame tests involve introducing a small amount of a metal ion into a flame to observe the characteristic colours emitted.
- The observed colours are due to the excitation of electrons to higher energy levels, followed by their return to lower levels, releasing energy in the form of light.
- The flame test is a qualitative analysis technique used to identify the presence of specific metal ions in a sample based on the colour of the flame they produce.

<u>Flame Test Procedure for Different Metals:</u>

Materials:
- Metal salts of lithium (Li), sodium (Na), potassium (K), calcium (Ca), barium (Ba), copper (Cu), strontium (Sr), rubidium (Rb), caesium (Cs)
- Nichrome wire loop
- Bunsen burner
- Safety goggles

<u>Procedure:</u>

Preparation:
- Gather the metal salts in separate containers.
- Clean the nichrome wire loop by dipping it into hydrochloric acid and then heating it in the Bunsen burner flame until no colour is visible.

Dipping the Loop:
- Dip the cleaned nichrome wire loop into the metal salt of interest. Ensure the loop is not contaminated with residues from previous tests.

Placing in Flame:
- Hold the wire loop in the hottest part of the Bunsen burner flame. Observe the colour of the flame.

Observation:
- Note the colour of the flame produced. Use a dark background to enhance visibility.

Safety Precautions:
- Wear safety goggles to protect your eyes from potential splashes or sparks.
- Work in a well-ventilated area.

Interpretation of Results:
- Match the observed flame colour with the characteristic flame colours associated with different metals.

Caution:
- Perform the flame test with caution, especially when dealing with potentially toxic metal salts.
- Dispose of waste materials according to safety guidelines.

Test for carbonates

The test for carbonates involves the reaction with an acid to produce carbon dioxide gas. Here's a general procedure:

Materials:

- Solid sample suspected to contain carbonate
- Dilute hydrochloric acid (HCl) or any other mineral acid
- Test tube
- Rubber stopper with delivery tube
- Lime water (calcium hydroxide solution)

Procedure:

Prepare the Sample:

- Place a small amount of the solid sample (suspected carbonate) in a test tube.

Add Acid:

- Add dilute hydrochloric acid (HCl) to the test tube containing the sample.

Observation 1 - Effervescence:

- If the sample contains carbonate ions, effervescence (bubbling) will occur. This is due to the release of carbon dioxide gas.

Collect Gas:

- Insert a rubber stopper with a delivery tube into the test tube. Allow the evolved gas (carbon dioxide) to pass through the delivery tube.

Observation 2 - Limewater Test:

- Bubble the evolved gas through lime water.
- If carbon dioxide is present, the lime water will turn milky or cloudy. This is because carbon dioxide reacts with lime water to form calcium carbonate, which is only slightly soluble and precipitates.

Interpretation of Results:

- Effervescence (bubbling) during the addition of acid indicates the presence of carbonate ions.
- Cloudiness or milkiness in lime water confirms the presence of carbon dioxide.

Caution:

- Perform the test in a well-ventilated area.
- Handle acids with care and wear appropriate safety equipment.

Test for Sulphate ions:

To test for the presence of sulfate ions (SO_4^{2-}), a barium chloride ($BaCl_2$) solution can be used. When barium chloride reacts with sulfate ions, it creates a white precipitate of barium sulfate ($BaSO_4$).

$$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4(s)$$

<u>False result</u>

Hydrochloric acid is necessary to react with carbonate impurities that are frequently found in salts. These impurities can cause a white barium carbonate precipitate and lead to inaccurate results. Sulfuric acid cannot be used because it contains sulfate ions, which could result in a false positive outcome.

Test for Positive ions (NH_4^+)

Testing for cations (positive ions) involves warming the substance with NaOH(aq) to test for ammonium ion NH_4^+, which forms NH_3.

Ammonia gas can be identified by its pungent smell or by turning red litmus paper blue.

Exam Questions

Which statement is **not** explained by hydrogen bonding?

- ☒ **A** all Group 1 hydroxides are soluble in water
- ☒ **B** many simple alcohols are soluble in water
- ☒ **C** the density of ice is less than the density of liquid water at 0 °C
- ☒ **D** the melting temperature of water is abnormally high

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A student made the following statements about trends going **down** Group 2.
Which statement is correct?

- ☒ **A** the thermal stability of the nitrates decreases
- ☒ **B** the thermal stability of the carbonates decreases
- ☒ **C** the solubility of hydroxides increases
- ☒ **D** the solubility of sulfates increases

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