
Inorganic Chemistry

Writing (Net Ionic) Equations for Various Reactions

On the AP examination you will encounter a question in which you will be required to write *net ionic equations* for various reactions. In past years, students have been required to choose 5 of 8 reactions. Some of the reactions you will undoubtedly recognize; others you will not! Hopefully, at least of them will be familiar to you!

On the examination you will be given reactions that actually occur - you will not be required to figure out whether or not the reaction actually works, as you did in 'Introductory' Chemistry. You will be required to write equations in *ionic form* - i.e. for **ionic compounds** you write the **anions (-)** and **cations (+)** as separate entities. You also omit the *spectator ions* - those ions that do not take part in the reaction. You wrote several different types of net ionic equations in 'INTRODUCTORY' Chemistry. On this handout, an attempt has been made to summarize the most important types of reactions. **It will not be possible, or useful, for you to learn dozens of specific reactions. Learn the general reactions, and figure out the specifics on the examination!**

FOR THE EXAMINATION YOU WILL NEED TO KNOW:

Formulae For Common Ions

Refer to the *Table of Ionic Charges* - hopefully all the *polyatomic ions* that you will encounter on the examination are on this chart.

Refer also to the *Table of Standard Reduction Potentials*, which you will be given on the examination.

Formulae For Common Acids

Again, refer to the *Table of Ionic Charges*

General Solubility Rules

Refer to the *Solubility Table* - you will need to know the general "trends".

ON THIS HAND-OUT REFERENCE IS MADE TO:

ELEMENTS - both **metals** and **non-metals**

ACIDS - both **dilute** and **concentrated**

BASES - **soluble metal hydroxides** (NaOH and KOH being the most common, by far)

- **soluble carbonates** (Na_2CO_3 and K_2CO_3 being the most common)

- an aqueous solution of **ammonia**, $\text{NH}_3(\text{aq})$

SALTS - **ionic** compounds that are neither acids nor bases. In general, **MX**, where **M** is a **metal** and **X** is an **anion** such as Cl^- , SO_4^{2-} , PO_4^{3-}

WATER

OXIDES

PEROXIDES

COMBUSTION REACTIONS

REDOX REACTIONS

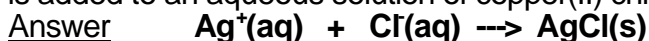
Precipitation (Double Displacement) Reactions

If you have **two salts** reacting, and a reaction occurs, the chances are excellent that it will be a *precipitation (double displacement)* reaction.

The precipitate will, obviously, be the insoluble salt.

Example

Write the *net ionic* equation for the reaction that occurs when an aqueous solution of silver nitrate is added to an aqueous solution of copper(II) chloride.



Refer back to the hand-out on *precipitation reactions* which you were given in Term 2.

There could be **two** precipitates, as is the case when an aqueous solution of copper(II) sulphate is added to an aqueous solution of barium hydroxide (**Both** copper(II) hydroxide **and** barium sulphate are **insoluble**)



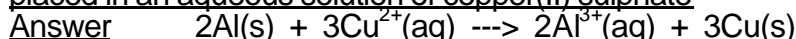
Single Displacement Reactions

In 'INTRODUCTORY' Chemistry, you did two "types" of *single displacement* reactions:

METAL ATOM + METAL ION

Example

Write the *net ionic* equation for the *single displacement* reaction that occurs when aluminum is placed in an aqueous solution of copper(II) sulphate

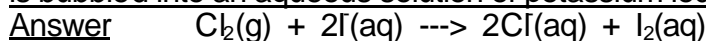


Use the E^o table provided on the examination to assist you with these equations!

HALOGEN MOLECULE + HALIDE ION

Example

Write the *net ionic* equation for the *single displacement* reaction that occurs when chlorine gas is bubbled into an aqueous solution of potassium iodide



Again, use the E^o table provided on the examination to assist you with these equations!

Reactions of Metals

WITH NON-METALS

Most metals can be made to react with most common non-metals (excluding the Noble gases); heat is usually required, however.

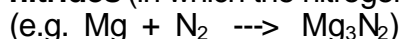
The reaction is usually a *synthesis*: METAL + NON-METAL → SALT

With oxygen oxide is formed

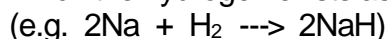
With halogens metal **halide** is formed

(When a metal with variable oxidation numbers, such as iron which exists both as Fe²⁺ and Fe³⁺, reacts with such things as oxygen and halogens, the compounds formed tend to be the ones in which the metal has the **higher** oxidation number. Thus, iron forms FeCl₃ when it reacts with chlorine, and Fe₃O₄ when it reacts with oxygen).

With nitrogen difficult to accomplish, but alkali metals and Mg from IIA form **nitrides** (in which the nitrogen is N³⁻ ion)



With hydrogen Group IA and IIA elements react with hydrogen to form metal **hydrides** (in which the hydrogen exists as H⁻ ion)



WITH WATER

If the metal reacts with pure water, the reaction can be summarized as:



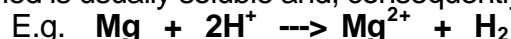
(Remember that *alkali metal* hydroxides are **soluble**, and the M^+ and OH^- ions will be written separately) E.g. $2K + 2H_2O \rightarrow 2K^+ + 2OH^- + H_2$

WITH DILUTE SOLUTIONS OF STRONG ACIDS (excluding HNO₃)

If a metal reacts with a dilute solution of acids such as HCl and H₂SO₄, the reaction that occurs can be summarized as:



(The salt formed is usually soluble and, consequently, the ions are written separately)

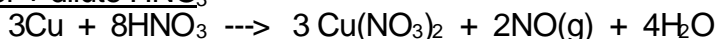


WITH DILUTE NITRIC ACID

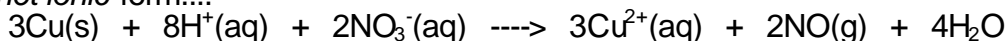
HNO₃ is a powerful oxidizing agent, and this is why it is usually employed in reactions.

Some examples you should know are:

Copper + dilute HNO₃



Or, in *net ionic* form....



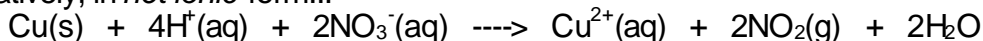
WITH CONCENTRATED ACIDS

There are, obviously, numerous different reactions involving concentrated solutions of strong acids. Unfortunately, they do not follow any general patterns, and are difficult to summarise. **You should, however, be aware of the following:**

Copper + concentrated HNO₃



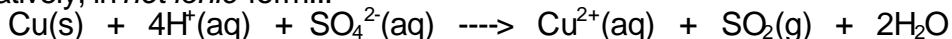
Alternatively, in *net ionic* form...



Copper + concentrated sulphuric acid



Alternatively, in *net ionic* form...



Reactions of Non-metals

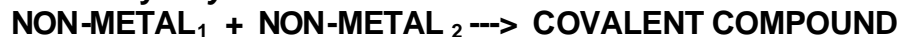
WITH METALS

This has already been discussed under "metals".

WITH OTHER NON-METALS

Most non-metals can be made to react with other non-metals (excluding the Noble gases); heat is usually required, however.

The reaction is usually a *synthesis*:



With oxygen

covalent **oxide** is formed (S forms SO₂, P forms P₄O₁₀)

With halogens

covalent **halide** is formed

With hydrogen

Covalent **hydrides** are formed. E.g. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

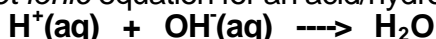
Reactions of Acids

NEUTRALIZATION REACTIONS

Reaction with Hydroxides

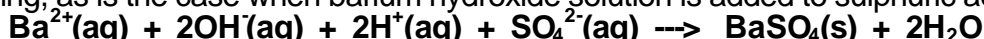


The *net ionic* equation for an acid/hydroxide neutralization usually come down to:



The salt formed is usually **soluble**, and the ions involved are, therefore, *spectator ions*.

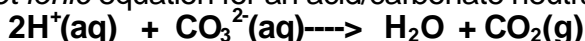
It is, however, possible to have **both** a *neutralization* reaction **and** a *precipitation* reaction occurring, as is the case when barium hydroxide solution is added to sulphuric acid:



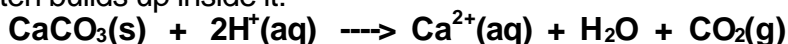
Reaction with Carbonates



The *net ionic* equation for an acid/carbonate neutralization very often comes down to:



This is the equation of the carbonate used is in aqueous solution. It could, however, be in solid form, as is the case if a (weak) acid is added to a kettle to dissolve the solid **scale** (CaCO_3) that often builds up inside it:

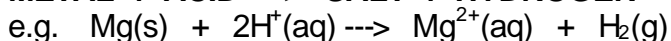


Also check to see if the salt formed is *soluble* or *insoluble*. Both of the above equations assume that the salt formed is **soluble**.

Reaction with Metals

METALS WITH DILUTE SOLUTIONS OF STRONG ACIDS (excluding HNO_3)

If a metal reacts with a dilute solution of acids such as HCl and H_2SO_4 , the reaction that occurs can be summarised as:



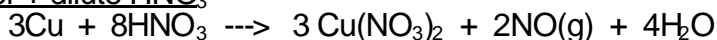
(The salt formed is usually soluble and, consequently, the ions are written separately)

METALS WITH DILUTE NITRIC ACID

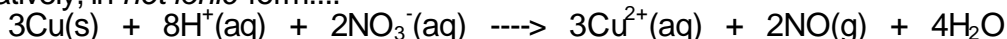
HNO_3 is a powerful oxidising agent, and this is why it is usually employed in reactions.

Some examples you should know are:

Copper + dilute HNO_3



Alternatively, in *net ionic* form....



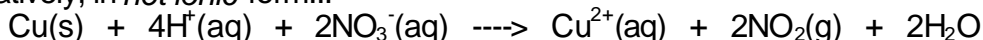
METALS WITH CONCENTRATED ACIDS

There are, obviously, numerous different reactions involving concentrated solutions of strong acids. Unfortunately, they do not follow any general patterns, and are difficult to summarise. **You should, however, be aware of the following:**

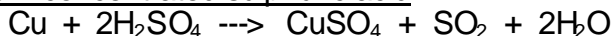
Copper + concentrated HNO_3



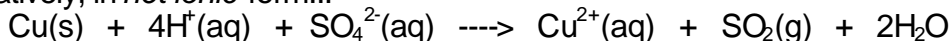
Alternatively, in *net ionic* form...



Copper + concentrated sulphuric acid



Alternatively, in *net ionic* form...



Reactions with Metallic Oxides

Since ionic oxides are **basic**, a **neutralization** reaction will occur when an ionic oxide reacts with a base (**salt** and **water** being formed).

Example

Write the *net ionic equation* for the reaction of solid magnesium oxide with dilute sulphuric acid

Answer: $\text{MgO(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{H}_2\text{O}$

Miscellaneous Reactions Involving Acids

Preparation of chlorine gas

Conc. HCl + MnO₂: $\text{MnO}_2(\text{s}) + 4\text{HCl} \rightarrow \text{MnCl}_2(\text{aq}) + \text{Cl}_2 + 2\text{H}_2\text{O}$

Or, in *net-ionic* form..... $\text{MnO}_2 + 4\text{H}^+ + 2\text{Cl}^- \rightarrow \text{Mn}^{2+} + \text{Cl}_2 + 2\text{H}_2\text{O}$

Conc. HCl + KMnO₄: $2\text{KMnO}_4(\text{s}) + 16\text{HCl} \rightarrow 2\text{KCl}(\text{aq}) + 2\text{MnCl}_2(\text{aq}) + 5\text{Cl}_2 + 8\text{H}_2\text{O}$

Or, in *net ionic* form..... $2\text{KMnO}_4 + 16\text{H}^+ + 10\text{Cl}^- \rightarrow 2\text{K}^+ + 2\text{Mn}^{2+} + 5\text{Cl}_2 + 8\text{H}_2\text{O}$

Reactions of Hydroxides

REACTION WITH ACIDS

ACID + HYDROXIDE \rightarrow SALT + WATER

The *net ionic* equation for an acid/hydroxide neutralization usually come down to:

$\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}$

The salt formed is usually **soluble**, and the ions involved are, therefore, *spectator ions*.

It is, however, possible to have **both** a *neutralization* reaction **and** a *precipitation* reaction occurring, as is the case when barium hydroxide solution is added to sulphuric acid:

$\text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}$

REACTION WITH NON-METALLIC (COVALENT) OXIDES AND WITH AMPHOTERIC OXIDES

Since covalent oxides are **acidic**, a **neutralization** reaction will occur when a covalent oxide reacts with a base (**salt** and **water** being formed).

Example

Write the *net ionic equation* for the reaction of sulphur dioxide gas with sodium hydroxide solution

Answer: $\text{SO}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}$

MISCELLANEOUS REACTIONS INVOLVING NaOH AND KOH

KOH(aq) and Hydrogen Sulphide (H₂S) gas

Hydrogen Sulphide is a *weak acid*, therefore **neutralization** occurs

$\text{H}_2\text{S}(\text{g}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{S}^{2-}(\text{aq}) + 2\text{H}_2\text{O}$ (K^+ ions are *spectator*)

Cl₂(g) and cold, dilute NaOH

$\text{Cl}_2(\text{g}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{OCl}^-(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}$

Reactions of Oxides

You have already learned that **metals** tend to form **ionic oxides** that are **basic**; whereas **non-metals** tend to form **covalent oxides** that are **acidic**. Aluminum, chromium and zinc form **amphoteric oxides**.

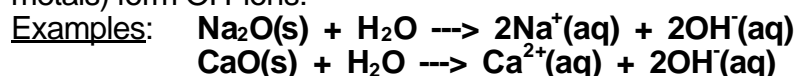
REACTIONS OF IONIC (BASIC) OXIDES

Decomposition

Most ionic oxides are very stable, but the oxides of metals low in the Activity Series do decompose on heating. For example: $2\text{Ag}_2\text{O} \rightarrow 4\text{Ag} + \text{O}_2$

Reaction With Water

Most ionic oxides are insoluble in water, but those that do dissolve (oxides of Group IA & IIA metals) form OH⁻ ions:

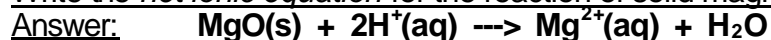


Reaction With Acids

Since ionic oxides are **basic**, a **neutralization** reaction will occur when an ionic oxide reacts with an acid (**salt** and **water** being formed).

Example

Write the net ionic equation for the reaction of solid magnesium oxide with dilute sulphuric acid



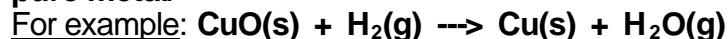
Reaction With Carbon

When the oxide is heated with carbon, the carbon will often **reduce** the oxide to the **pure metal**. This is often the way a metal is extracted from its ore.



Reaction With Hydrogen

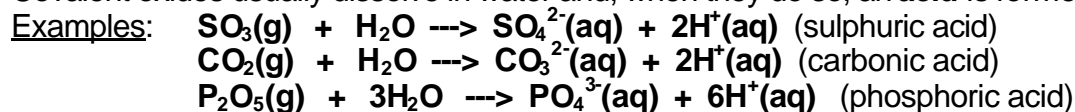
When the oxide is heated with hydrogen gas, the hydrogen will often **reduce** the oxide to the **pure metal**.



REACTIONS OF COVALENT (ACIDIC) OXIDES

Reaction With Water

Covalent oxides usually dissolve in water and, when they do so, an **acid** is formed.



Reaction With Bases

Since covalent oxides are **acidic**, a **neutralization** reaction will occur when a covalent oxide reacts with a base (**salt** and **water** being formed).

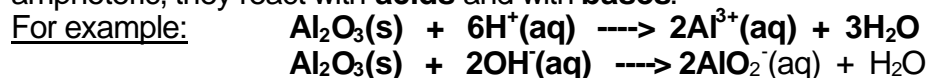
Example

Write the net ionic equation for the reaction of sulphur dioxide gas with sodium hydroxide solution



REACTIONS OF AMPHOTERIC OXIDES

Amphoteric oxides (such as Al₂O₃, Cr₂O₃ and ZnO) do not dissolve in water but, since they are amphoteric, they react with **acids** and with **bases**.

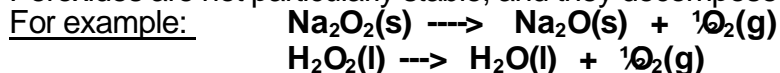


Reactions of Peroxides

A **peroxide** can be distinguished from the “normal” oxide by the fact that it contains more oxygen than “normal”. Peroxides actually contain the **O-O bond**.

Decomposition

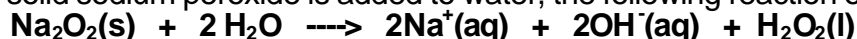
Peroxides are not particularly stable, and they decompose when heated gently.



Note that the decomposition of H_2O_2 can be **catalysed** by the addition of powdered **solid manganese dioxide (MnO_2)** or **KMnO_4** .

Reaction With Water

When solid sodium peroxide is added to water, the following reaction occurs:



Reactions of Water

REACTION WITH METALS

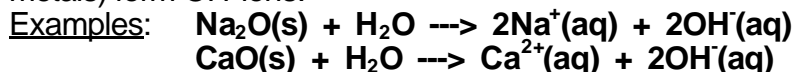
If the metal reacts with pure water, the reaction can be summarised as:



(Remember that *alkali metal* hydroxides are **soluble**, and the M^+ and OH^- ions will be written separately) E.g. $\text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca}^{2+} + 2\text{OH}^- + \text{H}_2$

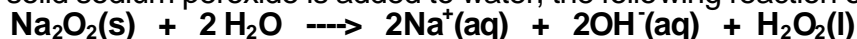
REACTION OF METAL OXIDES

Most ionic oxides are insoluble in water, but those that do dissolve (oxides of Group IA & IIA metals) form OH^- ions:



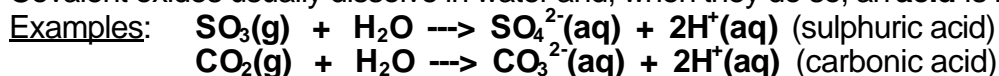
REACTION OF METAL PEROXIDES

When solid sodium peroxide is added to water, the following reaction occurs:



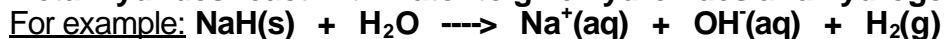
REACTION OF NON-METAL OXIDES

Covalent oxides usually dissolve in water and, when they do so, an **acid** is formed.



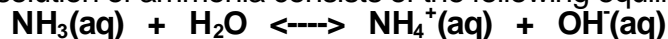
REACTION OF METAL HYDRIDES

Metal hydrides react with water to give hydroxides and hydrogen gas.



Reactions of Ammonia

An aqueous solution of ammonia consists of the following equilibrium mixture:

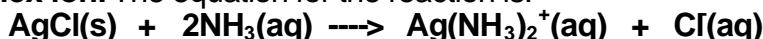


AMMONIA IS A WEAK BASE, SINCE IT CONTAINS THE OH⁻ ION

Ammonia will undergo the "hydroxide neutralization reactions" mentioned earlier, but remember it is a **weak** base.

COMPLEX ION FORMATION

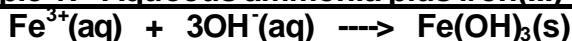
In Term 1 you found that $\text{NH}_3(\text{aq})$ would dissolve a white ppt. of $\text{AgCl}(\text{s})$ due to the formation of a **complex ion**. The equation for the reaction is:



PRECIPITATION REACTIONS

Since most **hydroxides are insoluble**, and an aqueous ammonia solution **contains OH⁻ ions**, numerous salts will form precipitates when mixed with aqueous ammonia solution.

Example 1: Aqueous ammonia plus iron(III) chloride solution



Example 2: Aqueous ammonia and zinc iodide solution



AMMONIA ACTING AS A LEWIS BASE

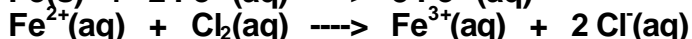
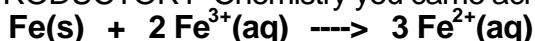
Remember: a **Lewis base** is an electron pair donor

Write the equation for the reaction that occurs between ammonia and boron fluoride

Oxidation-Reduction (Redox) Reactions

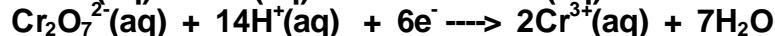
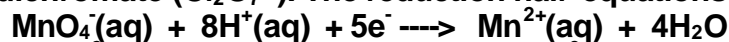
REACTIONS INVOLVING Fe(II) AND Fe(III) IONS

In 'INTRODUCTORY' Chemistry you came across the following reactions:



REACTIONS INVOLVING THE PERMANGANATE ION AND THE DICHROMATE ION

Two of the most common **oxidising agents** employed in the laboratory are **permanganate** (MnO_4^-) and **dichromate** ($\text{Cr}_2\text{O}_7^{2-}$). The **reduction half-equations** are as follows:



If you come across reactions involving either permanganate or dichromate, it is safe to assume that they will be **redox reactions**, and that the MnO_4^- or $\text{Cr}_2\text{O}_7^{2-}$ ions will be **reduced**.

Something else, obviously, will be **oxidised**. For example, Fe^{2+} could be oxidised to Fe^{3+} or Sn^{2+} could be oxidised to Sn^{4+}

Write net ionic equations for each of the following reactions

Iron(II) chloride solution reacting with an acidified solution of potassium permanganate

Tin(II) chloride solution reacting with an acidified solution of potassium dichromate

Combustion Reactions

ELEMENTS

Most elements undergo combustion to form **oxides**. The **synthesis** reaction can be summarised as: **ELEMENT + OXYGEN \rightarrow OXIDE**

ORGANIC COMPOUNDS

Most organic compounds undergo combustion. You have encountered these combustion reactions several times before.

The **hydrogen** in the organic compound becomes **water**.

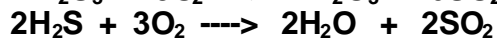
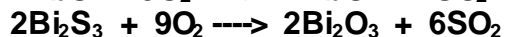
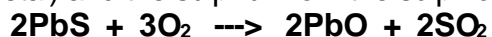
The **carbon** in the organic compound becomes **CO₂** when the combustion is **complete**, and either **CO** or **C** when the combustion is **incomplete**.

The **sulphur**, if it exists, becomes **SO₂(g)**.

The **nitrogen**, if it exists, becomes

METALLIC AND NON-METALLIC SULPHIDES

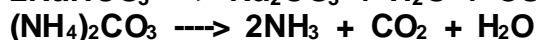
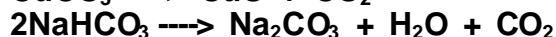
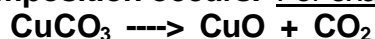
When a **metallic (or non-metallic) sulphide** is strongly heated in oxygen, both the metal (or non-metal) and the sulphur from the sulphide are **oxidised**. For example:



Reactions with Carbonates and Bicarbonates

HEATING CARBONATES AND BICARBONATES

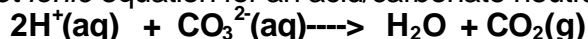
Decomposition occurs. For example:



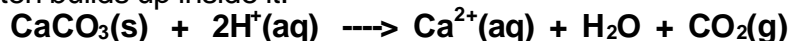
REACTIONS OF CARBONATES WITH DILUTE ACIDS



The *net ionic* equation for an acid/carbonate neutralization very often comes down to:



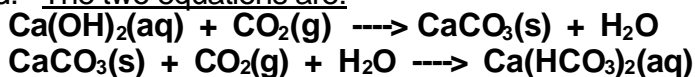
This is the equation of the carbonate used is in aqueous solution. It could, however, be in solid form, as is the case if a (weak) acid is added to a kettle to dissolve the solid **scale** (CaCO₃) that often builds up inside it:



Also check to see if the salt formed is *soluble* or *insoluble*. Both of the above equations assume that the salt formed is **soluble**.

LIMEWATER REACTIONS

Limewater is a solution of $\text{Ca}(\text{OH})_2$, and is employed as the usual test for CO_2 gas. When CO_2 is bubbled into limewater, it first turns "milky" due to the formation of a white ppt. of CaCO_3 . If you keep bubbling the CO_2 , however, the solution turns clear again since $\text{Ca}(\text{HCO}_3)_2(\text{aq})$ is formed. The two equations are:



Reactions for making some common gases in the laboratory

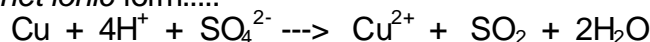
PREPARATION OF OXYGEN GAS BY HEATING POTASSIUM (OR SODIUM) CHLORATE



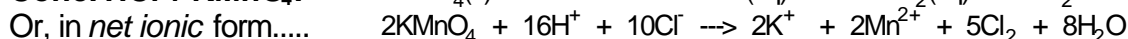
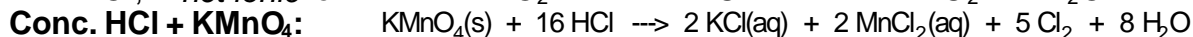
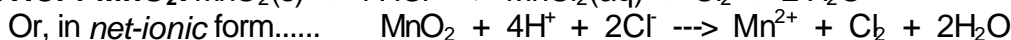
PREPARATION OF SULPHUR DIOXIDE BY REACTING COPPER WITH CONCENTRATED SULPHURIC ACID



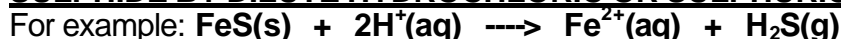
Or, written in *net ionic* form.....



PREPARATION OF CHLORINE GAS BY REACTING CONCENTRATED HYDROCHLORIC ACID EITHER WITH SOLID MANGANESE DIOXIDE OR WITH SOLID POTASSIUM PERMANGANATE

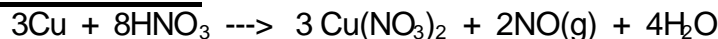


PREPARATION OF HYDROGEN SULPHIDE GAS BY REACTING A METALLIC SULPHIDE BY DILUTE HYDROCHLORIC OR SULPHURIC ACID

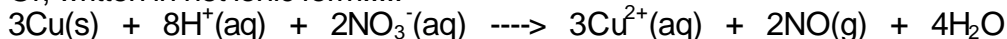


PREPARATION OF SOME OXIDES OF NITROGEN

Copper + dilute nitric acid



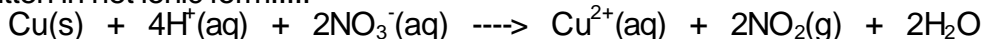
Or, written in net ionic form.....



Copper + concentrated nitric acid



Or, written in net ionic form.....

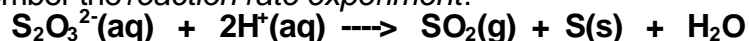


Miscellaneous Reactions



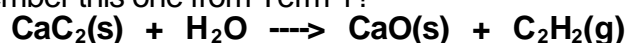
Sodium Thiosulphate and Hydrochloric Acid

Remember the *reaction rate experiment*?



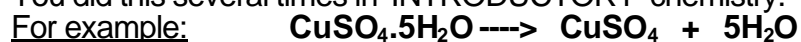
Preparing Acetylene

Remember this one from Term 1?



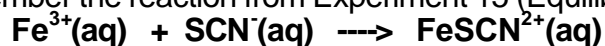
Heating a hydrate

You did this several times in 'INTRODUCTORY' chemistry.



Ammonium thiocyanate solution and iron(III) chloride solution

Remember the reaction from Experiment 15 (Equilibrium)?



Electrolysis Reactions

These will be covered in the final unit of the course. Do not despair, they are simple!!