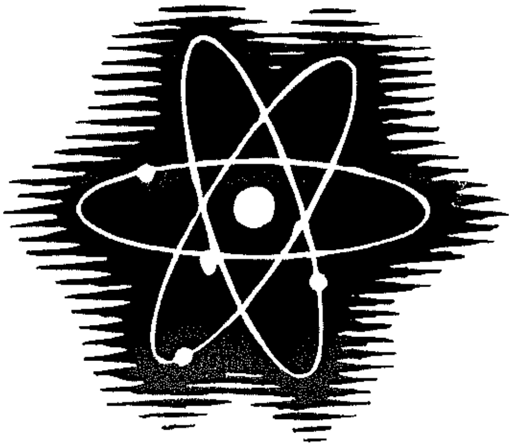
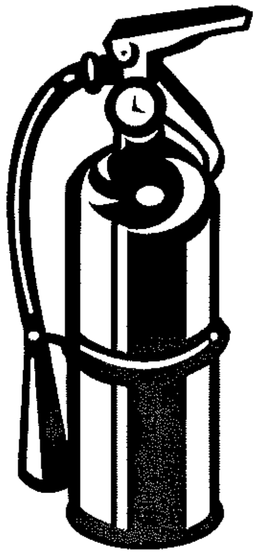
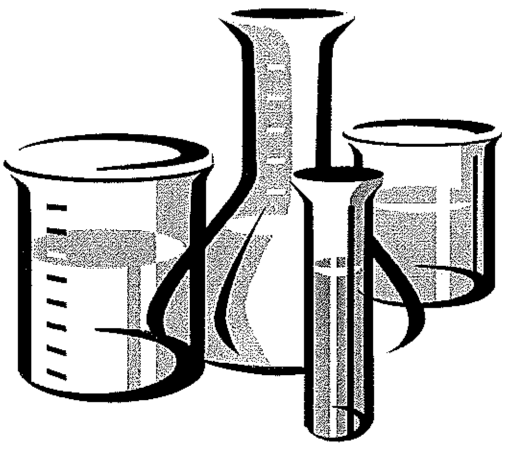
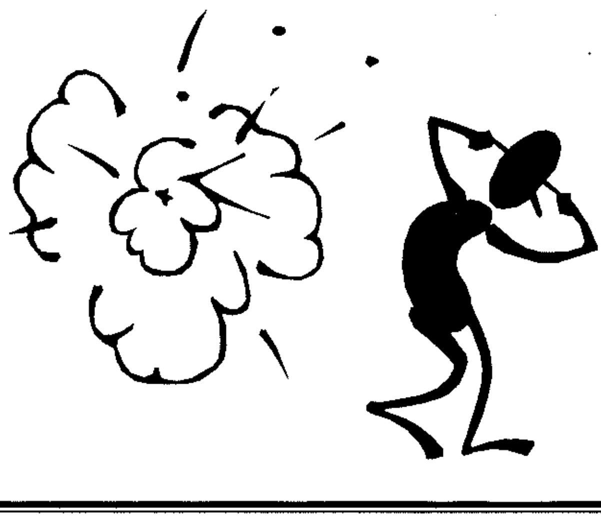




N4/N5
CHEMISTRY IN SOCIETY

FERTILISERS





WHAT I SHOULD KNOW : NATIONAL FIVE

5

• Describe using balanced formula equations how **ammonia** is used to make nitric acid • Describe using balanced formula equations how **nitric acid** is used to make nitrate fertilisers • State the **importance** of Nitrogen, Potassium and Phosphorus to plant growth • Calculate and compare the **percentage by mass** of N, P and K in different fertilisers • State that growing populations require more food and relate this to increased use of synthetic fertilisers • Research the **impact** synthetic fertilisers can have on the environment	Commercial production of nitrate fertilisers. Percentage mass compositions of fertilisers

• Show by experiment how to **produce** ammonia in the lab • Ammonia is a **soluble gas** that dissolves in water to produce an **alkali** • Explain why ammonia is difficult to produce by direct combination of nitrogen and hydrogen • Construct a balanced formula equation to show that the direct combination of nitrogen and hydrogen is a **reversible** reaction • Describe the industrial production of ammonia **Haber Process** • State **conditions** used in the Haber Process to maximise the yield of ammonia	Fertilisers Haber process to produce ammonia.

WHAT I SHOULD KNOW : NATIONAL FOUR

• Plants need the elements **nitrogen, phosphorus & potassium** to sustain healthy growth • State fertilisers, both natural & man-made, contain these elements and are used to **improve** plant growth • Work out by calculation, the **percentage composition** of an element in a fertiliser • Fertilisers can be **synthetic or natural** and give examples of natural fertilisers • Describe how to **prepare a fertiliser** by a neutralisation reaction • **Research & present** information about the impact of fertilisers on the environment	**Fertilisers** **Plant nutrients, and elements, natural and synthetic fertilisers**

Fertilisers - Why We Need Them N4

The world's population is increasing extremely rapidly.

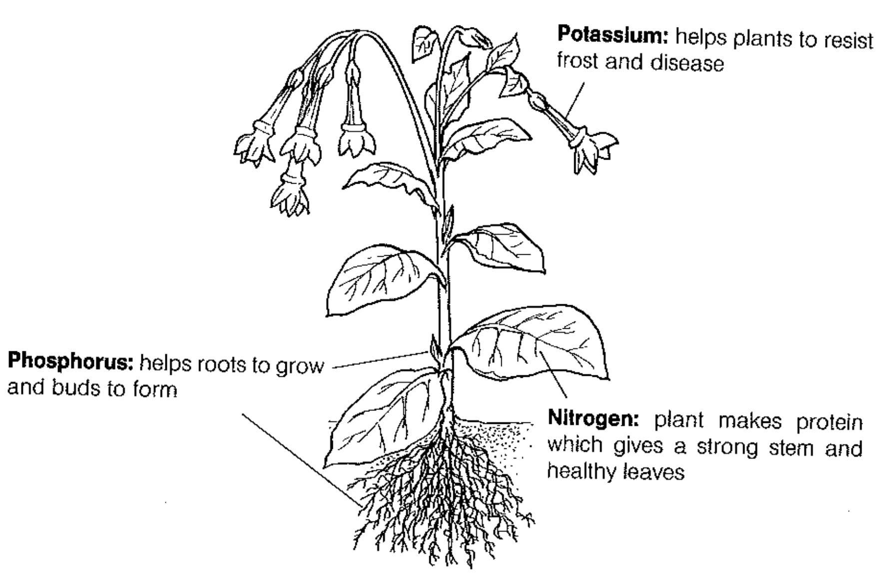
To continue to feed the world's population requires an increase in food production. All food starts with a plant. Plants need water and various nutrients to grow.

The main elements required by plants are shown in the table below.

Major Elements	Trace Elements
Carbon (from CO_2 in the air)	Iron
Hydrogen (from H_2O)	Manganese
Oxygen (from H_2O)	Copper
Nitrogen	Zinc
Phosphorous	Molybdenum
Sulphur	Boron
Potassium	Chlorine
Calcium	
Magnesium	

The elements (with the exception of C, H, O) come from the soil.

The 3 most important elements obtained from the soil are nitrogen, phosphorous and potassium.



Replacing Nitrogen N4

N, P and K are quickly removed from the soil by plants - they need to be replaced. For nitrogen this can happen naturally

Leguminous Plants

Legumes are plants like peas, beans and clover
They have nodules on their roots which contain nitrifying bacteria
These bacteria can take nitrogen from the air and make nitrates
This is known as fixing nitrogen
Because nitrates are soluble, the plant can absorb them and use the nitrogen to make protein

Compost

When plants die, they rot away
Ammonium salts containing nitrogen are released
Bacteria in the soil convert this to nitrates which are readily absorbed by the plant.

Manure

Animal waste is also rich in ammonium salts
When animals die they also release ammonium salts into the soil
These are converted into nitrates and absorbed by plants

Natural and Synthetic Fertilisers

Fertilisers are substances which replace the essential elements in the soil
They can be
-Natural like compost, manure or leguminous plants
-Synthetic which means manufactured

Synthetic Fertilisers N4

Synthetic or man-made fertilisers include ammonia and ammonium salts, potassium salts, nitrates and phosphates.
Compounds used as fertilisers must be **soluble** in water so that they can be absorbed by plants.
Complete the following table by writing the ion formulae and stating whether the salts would be soluble or insoluble:

Negative Ion / *Positive Ion*	**Sulphate**	**Nitrate**	**Phosphate**	**Chloride**
Ammonium				
Potassium				
Calcium				

Which salts in the table would make the best fertilisers?

___

Which salts would not be commonly used as fertilisers?

___

What are the **disadvantages** of using synthetic fertilisers?

___

___

___

Using nitrifying bacteria is a more economic and environmentally friendly method of adding nitrogen to soil because the nitrates are not washed away.

Fertilisers - What Is In Them? N4

Not all fertilisers contain the same amount of N, P and K

Name of Fertiliser	Used For	Proportion of		
		N	P	K
Fisons' Deep Feed				
Levington flowerrite				
B&Q liquid tomato feed				
ICI fertilizers				

A bag of NPK fertilizer. The percentages of N, P and K are indicated at the top of the bag—17%N, 17%P and 17%K

Calculating Percentage Composition N4/5

Many soluble compounds containing N, P and K could in theory be used as fertilisers. So how do you choose? If you require a nitrogen fertiliser then the percentage by mass of nitrogen in each compound can be compared. This can be calculated by following these steps:-

1. Work out the formula
2. Calculate the formula mass
3. Work out the % of N, P and K using the formula

$$\text{percentage}(\%)\ N = \frac{\text{mass of nitrogen in 1 mole}}{\text{mass of 1 mole}} \times 100$$

e.g. **What percentage nitrogen is found in ammonium phosphate?**

1. Work out the formula
   $(NH_4)_3PO_4$
2. Work out the formula mass
   $([14 + 1 + 1 + 1 + 1] \times 3) + 31 + (4 \times 16) = 149$
   mass of 1 mole = 149g
3. Work out the percentage of N present
   mass of nitrogen in 1 mole
   3 N present = 42g

$$\text{percentage } N = \frac{\text{mass of nitrogen in 1 mole}}{\text{mass of 1 mole}} \times 100$$

$$= 42 / 149 \times 100 = \underline{\mathbf{28.2\%}}$$

To work out the percentage of phosphorous in the same compound

3. Work out the percentage of P present
   mass of phosphorous in 1 mole
   1 P present = 31g

$$\text{percentage } P = \frac{\text{mass of phosphorous in 1 mole}}{\text{mass of 1 mole}} \times 100$$

$$= 31 / 149 \times 100 = \underline{\mathbf{20.8\%}}$$

Percentage Composition N4

1. Calculate the % N in the following compounds:

a) ammonia, NH_3

b) ammonium nitrate

c) sodium nitrate

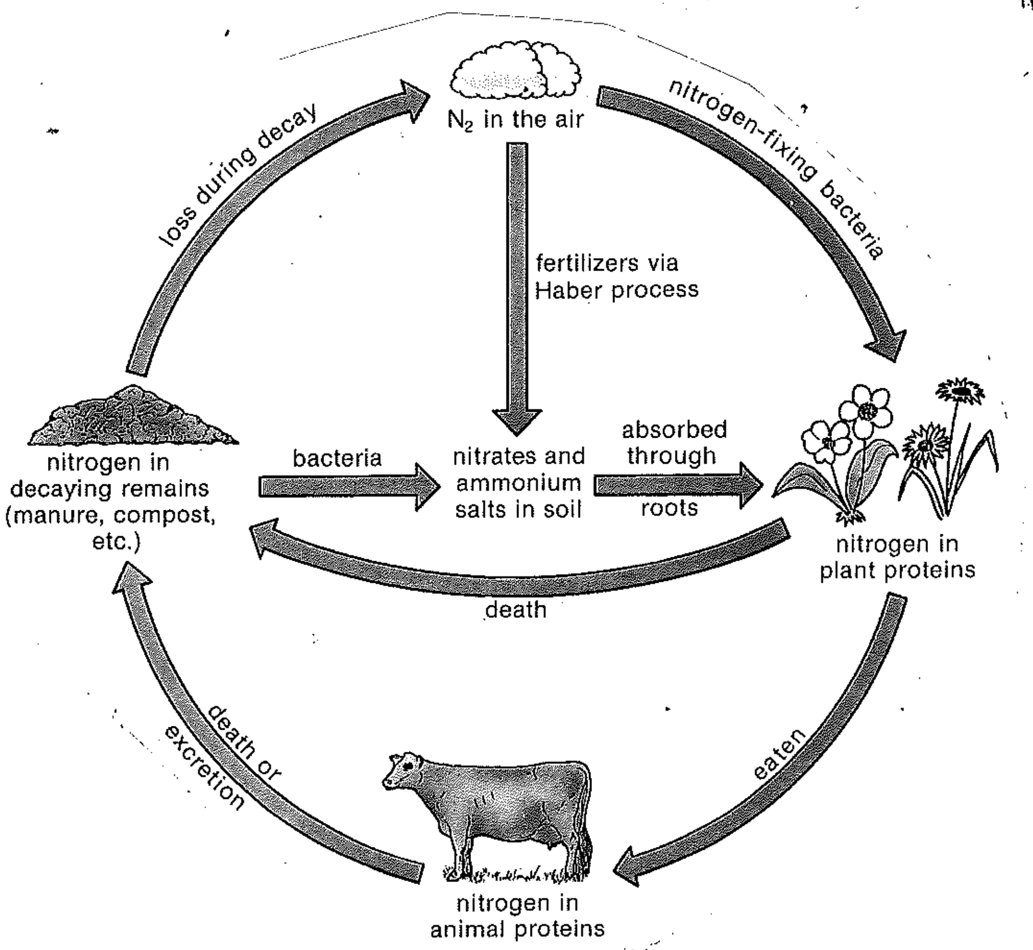
d) urea, $CO(NH_2)_2$

2. Calculate the % phosphorous in the following:

a) ammonium hydrogen phosphate, $(NH_4)_2HPO_4$

b) calcium superphosphate, $Ca(H_2PO_4)_2$

3. Calculate the % potassium in potassium nitrate



The above diagram summarises how nitrogen is added to and lost from the soil in <u>**THE NITROGEN CYCLE**</u>.

Removal:

i) plants are 'harvested' and taken away

ii) nitrates (containing nitrogen) are **soluble** and can be washed away.

iii) Denitrifying bacteria break up nitrogen compounds in the soil and release the nitrogen into the air.

Addition:

a) dead plants and animals decay with the nitrogen being returned to the soil (compost).

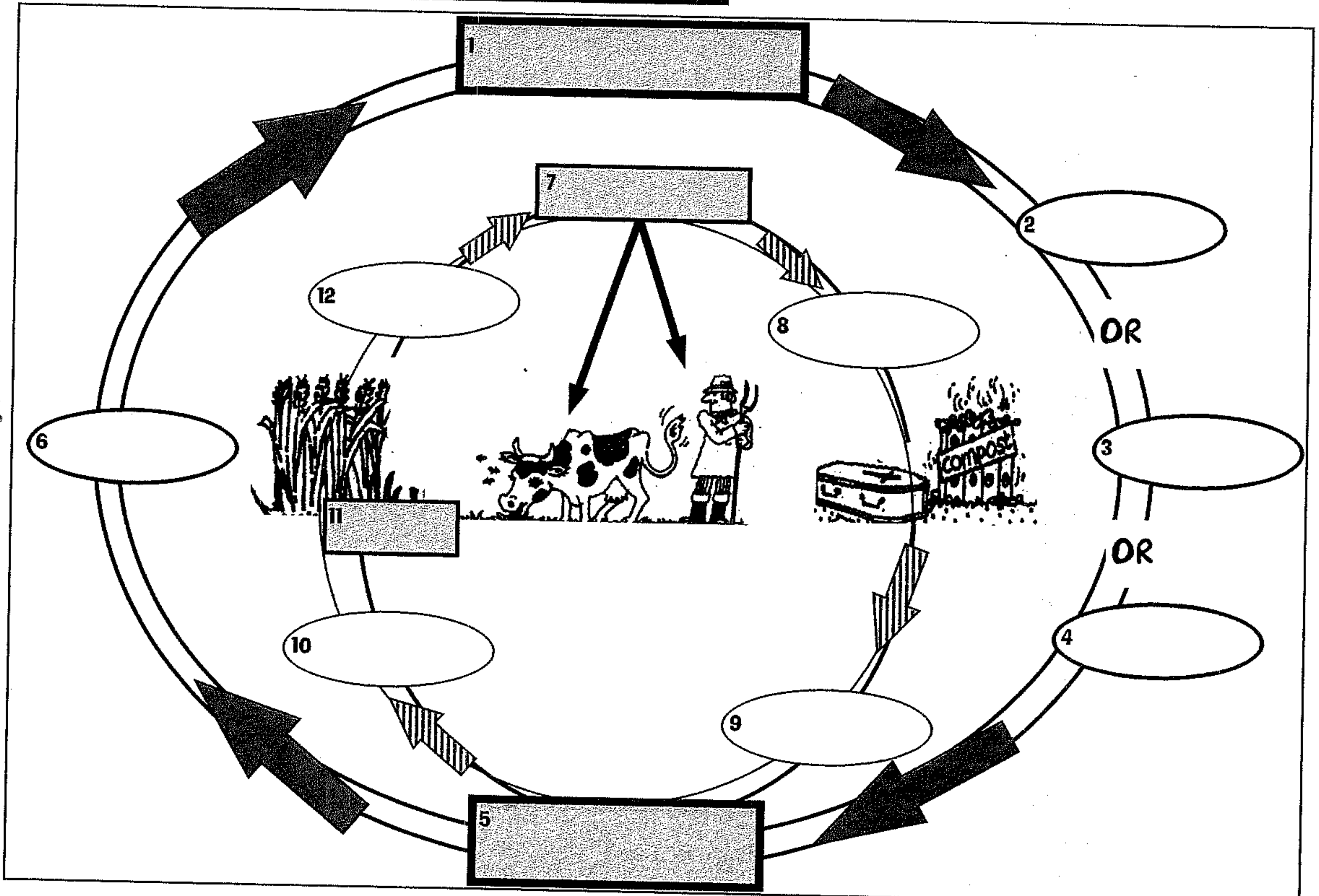
b) animal manure

c) Legumes (peas, beans and clover) have nodules on their roots able to take the nitrogen from the air and make it into soluble nitrates to be used by the plants.

d) The 'flash' during a lightning storm can cause the nitrogen and the oxygen in the air to combine. The nitrogen oxide formed dissolves in rain water and is able to enter the soil.

e) Artificial fertilisers.

Fertilisers 3 : The Nitrogen Cycle



lightning

denitrifying bacteria

death

decomposers and decay

uptake by roots

digestion in animals

fertilisers

nitrifying bacteria

NITRATE IN *PLANTS*

NITRATE IN *SOIL*

NITRATE IN *ANIMALS*

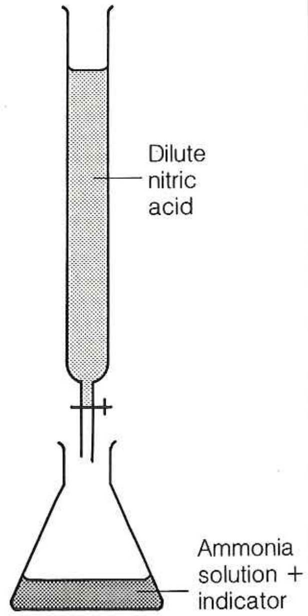
NITROGEN GAS IN THE *AIR*

Making Fertilisers N4

Many fertilisers contain the salt ammonium nitrate.

Which acid and alkali are required to form this salt?

_______________ and _______________



Ammonia as a base

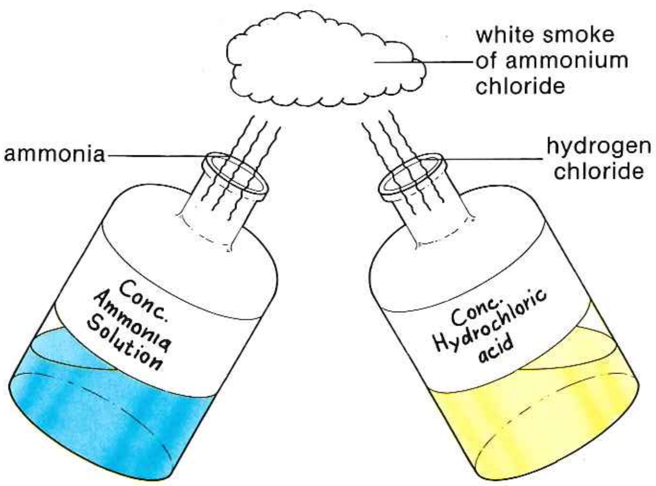
Ammonia acts as a base in many reactions. Ammonia molecules pick up H^+ ions to form ammonium ions. So, ammonia reacts with acids to form ammonium salts. This is how fertilizers, such as ammonium nitrate ('Nitram') and ammonium sulphate, are made from ammonia.

$$\underset{\text{ammonia}}{NH_3(g)} + \underset{\text{nitric acid}}{HNO_3(aq)} \rightarrow \underset{\text{ammonium nitrate}}{NH_4NO_3(aq)}$$

Ammonia also reacts with hydrogen chloride to form a white smoke. The white smoke is tiny particles of solid ammonium chloride suspended in the air. This reaction is sometimes used as a test for ammonia.

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

When ammonia dissolves in water, it acts as a base by taking H^+ ions from the water to form NH_4^+ ions.



The Haber Process N5

1. Name the two gases used to make ammonia.

   a) ____________________ b) ____________________

2. Where do we get these gases?

   a) ____________________ b) ____________________

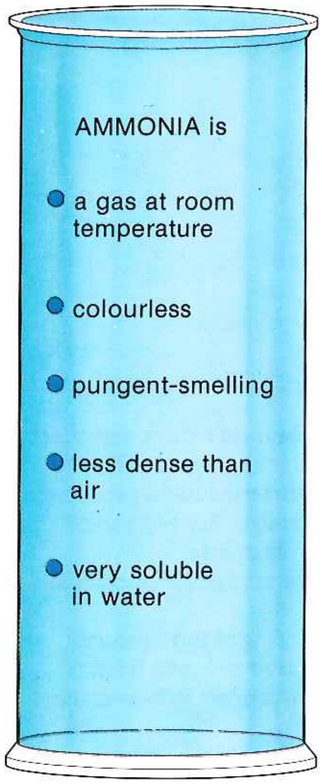
3. What temperature and pressure is used in the process?

   Temperature = __________ Pressure = ____________

4. Which catalyst is used in the process? ____________

   What does a catalyst do? ______________________

   ____



5. Write the word equation for the Haber Process

6. Write a balanced equation for the reaction

7. What does the $\rightleftharpoons$ sign mean?

   ____

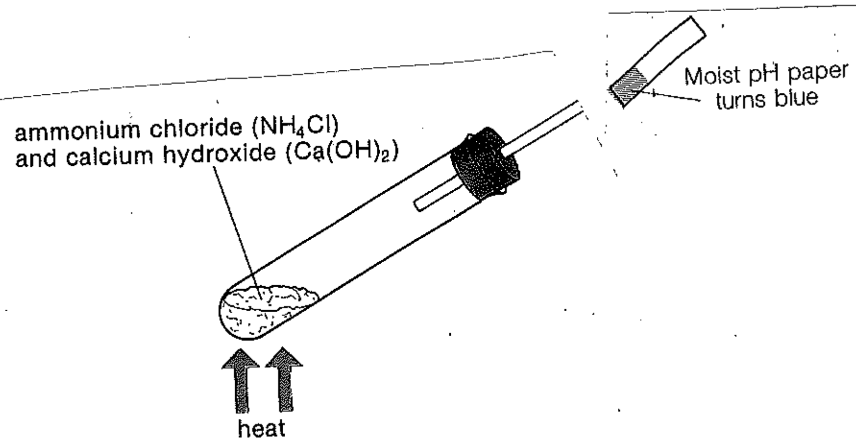
8. Why is a moderately high temperature used as a compromise in the reaction?

   ____

   ____

Making Ammonia in the Lab N5

Ammonia can be prepared in the lab using the apparatus shown below.



Properties of Ammonia

Appearance	
Smell	
Solubility	
pH	

Wet nappies give off ammonia gas!

Ammonia N5

Ammonia (NH_3) is used as a fertiliser and is used to make other fertilisers.

Ammonia is manufactured by the **Haber process**. In this process, nitrogen (from the air) and hydrogen (from methane obtained from natural gas) are passed over an iron catalyst at moderately high temperature (about 500 °C) and high pressure (about 200 atmospheres). The process is reversible and so the mixture is cooled so that the liquefied ammonia can be removed. The unreacted nitrogen and hydrogen are recycled back to the catalyst chamber.

$$N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2\,NH_{3(g)}$$

The Haber process is carried out at **moderately** high temperatures because:

a) the reaction is too **slow** at lower temperatures
b) the percentage conversion is too **low** at higher temperatures

All the nitrogen and hydrogen are **not** converted to ammonia because the reaction is **reversible.**

Properties of Ammonia

- Ammonia is a colourless gas with a sharp, unpleasant (pungent) smell
- Ammonia is very soluble in water producing an **alkaline** solution
  $NH_{3(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$
- Ammonia is the only common alkaline gas
- Ammonia can act as a neutraliser with acids, accepting hydrogen ions
  $NH_{3(aq)} + H^+{}_{(aq)} \rightarrow NH_4^+{}_{(aq)}$ (ammonium ion)
- Ammonia reacts with acids producing ammonium salts, e.g.
  $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$ (ammonium sulphate)
- Ammonia gas is given off when an **ammonium salt** is heated with an **alkali** such as sodium hydroxide
  $NH_4Cl + NaOH \rightarrow NaCl + H_2O + NH_3$

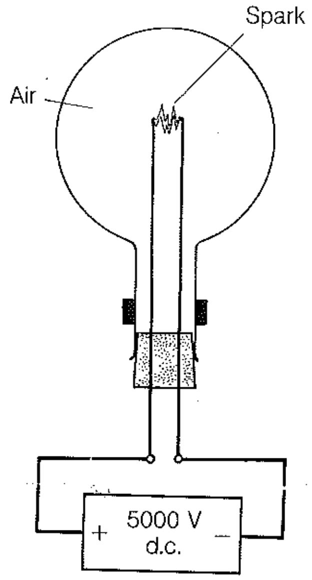
Nitric Acid N5

Nitrogen gas is not very reactive. For nitrogen and oxygen in the air to combine, a high energy spark is required. This spark can be provided by:

1. ______________________________

2. ______________________________

When nitrogen and oxygen combine in this way, nitrogen dioxide is produced. The equipment shown below can be used to make nitrogen dioxide in the lab.

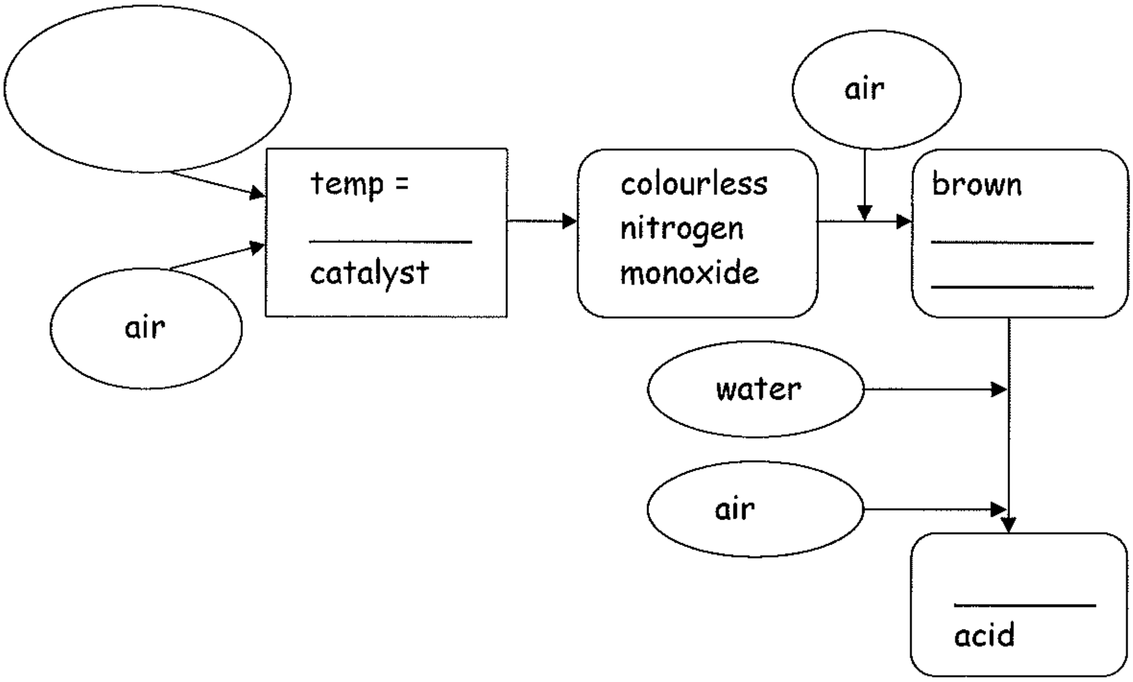


Appearance	
Smell	
Solubility	
pH	

The Ostwald Process N5

Nitrogen dioxide can be produced when a high energy spark is passed through air. This nitrogen dioxide dissolves in water to form nitric acid. However, this is not a economic way to make nitric acid industrially.

Nitric acid is made industrially through the Ostwald Process.



The word equations for the Ostwald Process are:

ammonia + oxygen → nitrogen monoxide + water

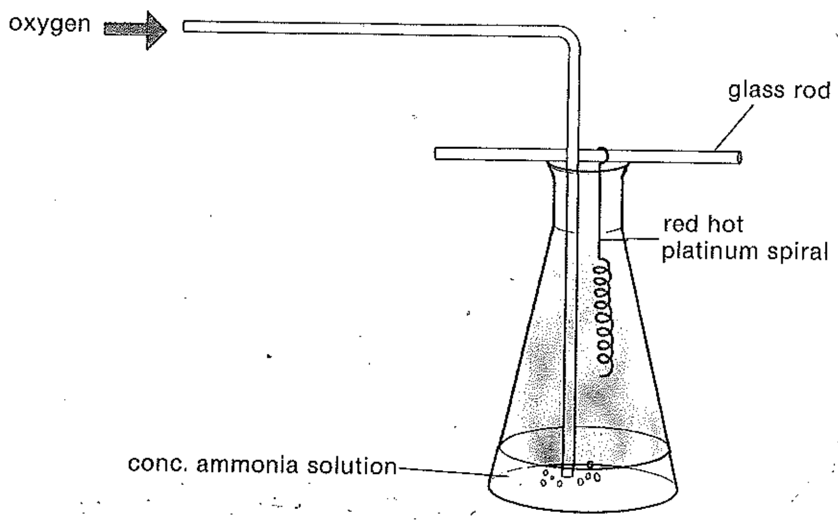
nitrogen monoxide + oxygen → nitrogen dioxide

nitrogen dioxide + oxygen + water → nitric acid

The Catalytic Oxidation of Ammonia

N5

The reaction between ammonia and oxygen to make nitrogen dioxide can be carried out in the laboratory using a hot platinum wire as catalyst.



Once the reaction has begun, it is no longer necessary to keep heating the catalyst. Why is this?

____

Explain why the reaction is carried out at a moderately high temperature

____

____