Extraction of Metals

How easy it is to get a metal out of its ore all comes down to the metal's position in the reactivity series.

More Reactive Metals are Harder to Get

1) A few unreactive metals like gold are found in the Earth as the metal itself, rather than as a compound.
2) But most metals need to be extracted from their ores using a chemical reaction.
3) More reactive metals, like sodium, are harder to extract — that's why it took longer to discover them.

Some Metals can be Extracted by Reduction with Carbon

1) Electrolysis (splitting with electricity) is one way of extracting a metal from its ore. The other common way is chemical reduction using carbon or carbon monoxide.
2) When an ore is reduced, oxygen is removed from it, e.g.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

iron(III) oxide + carbon monoxide → iron + carbon dioxide

3) The position of the metal in the reactivity series determines whether it can be extracted by reduction with carbon or carbon monoxide.

a) Metals higher than carbon in the reactivity series have to be extracted using electrolysis, which is expensive.

b) Metals below carbon in the reactivity series can be extracted by reduction using carbon. For example, iron oxide is reduced in a blast furnace to make iron.

This is because carbon can only take the oxygen away from metals which are less reactive than carbon itself is.

The Reactivity Series

Extracted using Electrolysis	Potassium	K	more
	Sodium	Na	reactive
	Calcium	Ca	
	Magnesium	Mg	
	Aluminium	Al	
	CARBON	C	
Extracted by reduction using carbon	Zinc	Zn	
	Iron	Fe	
	Tin	Sn	less
	Copper	Cu	reactive

Even primitive folk could find gold easy enough just by scrabbling about in streams, and then melt it into ingots and jewellery and statues of ABBA during their 1857 comeback tour, but coming up with a fully operational electrolysis plant to extract sodium metal from rock salt, complete with plastic yuccas in the foyer, just by paddling about a bit... unlikely.

A More Reactive Metal Displaces a Less Reactive Metal

1) More reactive metals react more strongly than less reactive metals.
2) This means that a metal can be extracted from its oxide by any more reactive metal. The more reactive metal bonds more strongly to the oxygen and pushes out the less reactive metal. E.g. tin could be extracted from tin oxide by more reactive iron.

   tin oxide + iron → iron oxide + tin

3) And if you put a reactive metal into the solution of a dissolved metal compound, the reactive metal will replace the less reactive metal in the compound. E.g. put an iron nail in a solution of copper sulfate and the more reactive iron will "kick out" the less reactive copper from the solution. You end up with iron sulfate solution and copper metal.

   copper sulfate + iron → iron sulfate + copper

4) If a piece of silver metal is put into a solution of copper sulfate, nothing happens. The more reactive metal (copper) is already in the solution.

Learn how metals are extracted — ore else...

Extracting metals isn't cheap. You have to pay for special equipment, energy and labour. Then there's the cost of getting the ore to the extraction plant. If there's a choice of extraction methods, a company always picks the cheapest, unless there's a good reason not to (like with copper). They're not extracting it for fun.

Relative Formula Mass

The biggest trouble with relative atomic mass and relative formula mass is that they sound so blood-curdling. Take a few deep breaths, and just enjoy, as the mists slowly clear...

Relative Atomic Mass, A_r — Easy Peasy

1) This is just a way of saying how heavy different atoms are compared with the mass of an atom of carbon-12. So carbon-12 has an A_r of exactly 12.
2) It turns out that the relative atomic mass A_r is nothing more than the mass number of the element (to the nearest whole number).
3) In the periodic table, the elements all have two numbers. The smaller one is the atomic number (how many protons it has). But the bigger one is the mass number (how many protons and neutrons it has), which is also the relative atomic mass. Easy peasy, I'd say.

Mass Number which is also Relative Atomic Mass

$^{4}_{2}He$ $^{12}_{6}C$ $^{35.5}_{17}Cl$

You may have noticed that the relative atomic mass of chlorine isn't a whole number like the others. That's because it has more than one stable isotope — it's all explained on page 58.

Helium has $A_r = 4$. Carbon has $A_r = 12$. Chlorine has $A_r = 35.5$.

Relative Formula Mass, M_r — Also Easy Peasy

If you have a compound like $MgCl_2$ then it has a relative formula mass, M_r, which is just all the relative atomic masses added together.

For $MgCl_2$ it would be:

$$MgCl_2$$

$$24 \quad + \quad (35.5 \times 2) \quad = \quad 95$$

So M_r for $MgCl_2$ is simply 95.

You can easily get A_r for any element from the periodic table (see inside front cover), but in a lot of questions they give you them anyway. I tell you what, since it's nearly Christmas I'll run through another example for you:

Question: Find the relative formula mass for calcium carbonate, $CaCO_3$, using the given data:
A_r for Ca = 40 A_r for C = 12 A_r for O = 16

ANSWER:

$$CaCO_3$$

$$40 \quad + \quad 12 \quad + \quad (16 \times 3) \quad = 100$$

So the relative formula mass for $CaCO_3$ is 100.

And that's all it is. A big fancy name like relative formula mass and all it means is "add up all the mass numbers". What a swizz, eh? You'd have thought it'd be something a bit juicier than that, wouldn't you. Still, that's life — it's all a big disappointment in the end. Sigh.

Numbers? — and you thought you were doing chemistry...

Learn the definitions of relative atomic mass and relative formula mass, then have a go at these:

1) Use the periodic table to find the relative atomic mass of these elements: Cu, K, Kr, Cl
2) Find the relative formula mass of: NaOH, Fe_2O_3, C_6H_{14}, $Mg(NO_3)_2$

Answers on page 124.

Isotopes and Relative Atomic Mass

Some elements have more than one isotope.

Isotopes are the Same Except for an Extra Neutron or Two

A favourite trick exam question: "Explain what is meant by the term isotope"
The trick is that it's impossible to explain what one isotope is. Nice of them that, isn't it.
You have to outsmart them and always start your answer "Isotopes are..." LEARN the definition:

Isotopes are: different atomic forms of the same element, which have the SAME number of PROTONS but DIFFERENT numbers of NEUTRONS.

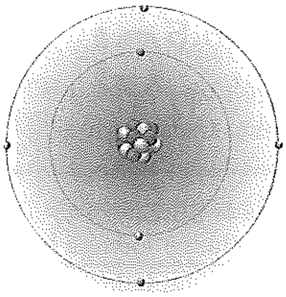
1) The upshot is: isotopes must have the same proton number but different mass numbers.
2) If they had different proton numbers, they'd be different elements altogether.
3) A very popular pair of isotopes are carbon-12 and carbon-14, used for carbon dating.

See page 4 for more about atomic structure.

Carbon-12

$^{12}_{6}C$

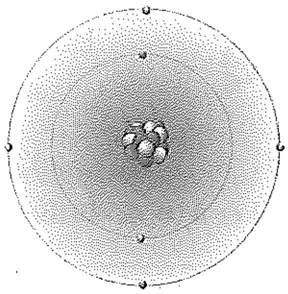
6 PROTONS
6 ELECTRONS
6 NEUTRONS



Carbon-14

$^{14}_{6}C$

6 PROTONS
6 ELECTRONS
8 NEUTRONS



The number of electrons decides the chemistry of the element. If the atomic number (that is, the number of protons) is the same, then the number of electrons must be the same, so the chemistry is the same. The different number of neutrons in the nucleus doesn't affect the chemical behaviour at all.

Relative Atomic Mass Takes All Stable Isotopes into Account

1) Relative atomic mass (A_r) uses the average mass of all the isotopes of an element. It has to allow for the relative mass of each isotope and its relative abundance.
2) Relative abundance just means how much there is of each isotope compared to the total amount of the element in the world. This can be a ratio, a fraction or a percentage and is easiest to see with an example:

element	relative mass of isotope	relative abundance
chlorine	35	3
	37	1

This means that there are 2 isotopes of chlorine. One has a relative mass of 35 (^{35}Cl) and the other 37 (^{37}Cl).
The relative abundances show that there are 3 atoms of ^{35}Cl to every 1 of ^{37}Cl.

1) First, multiply the mass of each isotope by its relative abundance.
2) Add those together.
3) Divide by the sum of the relative abundances.

$$A_r = \frac{(35 \times 3) + (37 \times 1)}{3 + 1} = 35.5$$

Relative atomic masses don't usually come out as whole numbers or easy decimals, but they're often rounded to the nearest 0.5 in periodic tables (see p57).

Will this be in your exam? — isotope so...

Some isotopes are unstable. That means they don't stay as they are forever, but change (decay) into other elements. When they do this, they release nuclear radiation.

Percentage Mass and Empirical Formulas

Although relative atomic mass and relative formula mass are easy enough, it can get just a tad trickier when you start getting into other calculations which use them. It depends on how good your maths is basically, because it's all to do with ratios and percentages.

Calculating % Mass of an Element in a Compound

This is actually dead easy — so long as you've learnt this formula:

$$\text{Percentage mass OF AN ELEMENT IN A COMPOUND} = \frac{A_r \times \text{No. of atoms (of that element)}}{M_r \text{ (of whole compound)}} \times 100$$

If you don't learn the formula then you'd better be pretty smart — or you'll struggle.

Find the percentage mass of nitrogen in ammonium sulfate fertiliser, $(NH_4)_2SO_4$, using the following:
A_r for H = 1 A_r for N = 14 A_r for O = 16 A_r for S = 32

ANSWER: M_r of $(NH_4)_2SO_4 = 2 \times [14 + (1 \times 4)] + 32 + (16 \times 4) = 132$

Now use the formula:

$$\text{Percentage mass} = \frac{A_r \times n}{M_r} \times 100 = \frac{14 \times 2}{132} \times 100 = 21.2\%$$

So there you have it. Nitrogen represents 21.2% of the mass of ammonium sulfate.

Finding the Empirical Formula (from Masses or Percentages)

This also sounds a lot worse than it really is. Try this for a nice simple stepwise method:

1) List all the elements in the compound (there's usually only two or three)
2) Underneath them, write their experimental masses or percentages.
3) Divide each mass or percentage by the A_r for that particular element.
4) Turn the numbers you get into a nice simple ratio by multiplying and/or dividing them by well-chosen numbers.
5) Get the ratio in its simplest form, and that tells you the empirical formula of the compound.

Example: Find the empirical formula of the iron oxide produced when 44.8 g of iron react with 19.2 g of oxygen. (A_r for iron = 56, A_r for oxygen = 16)

Method:

1) List the two elements:	Fe	O
2) Write in the experimental masses:	44.8	19.2
3) Divide by the A_r for each element:	$44.8/56 = 0.8$	$19.2/16 = 1.2$
4) Multiply by 10...	8	12
...then divide by 4:	2	3

5) So the simplest formula is 2 atoms of Fe to 3 atoms of O, i.e. Fe_2O_3. And that's it done.

You need to realise (for the exam) that this empirical method (i.e. based on experiment) is the only way of finding out the formula of a compound. Rust is iron oxide, sure, but is it FeO, or Fe_2O_3? Only an experiment to determine the empirical formula will tell you for certain.

With this empirical formula I can rule the world! — mwa ha ha ha...

Make sure you learn the formula and the five rules in the red box. Then try these: Answers on page 124

1) Find the percentage mass of oxygen in each of these: a) Fe_2O_3 b) H_2O c) $CaCO_3$ d) H_2SO_4.
2) Find the empirical formula of the compound formed from 2.4 g of carbon and 0.8 g of hydrogen.

Calculating Masses in Reactions

These can be a bit scary too, but chill out, little trembling one — just relax and enjoy.

The Three Important Steps — Not to be Missed...

(Miss one out and it'll all go horribly wrong, believe me.)

1) Write out the balanced equation
2) Work out M_r — just for the two bits you want
3) Apply the rule: Divide to get one, then multiply to get all
   (But you have to apply this first to the substance they give information about, and then the other one!)

Don't worry — these steps should all make sense when you look at the example below.

Example: What mass of magnesium oxide is produced when 60 g of magnesium is burned in air?

Answer:

1) Write out the balanced equation: $2Mg + O_2 \rightarrow 2MgO$
2) Work out the relative formula masses:
   (don't do the oxygen — you don't need it) $2 \times 24 \rightarrow 2 \times (24+16)$
   $48 \rightarrow 80$
3) Apply the rule: Divide to get one, then multiply to get all
   The two numbers, 48 and 80, tell us that 48 g of Mg react to give 80 g of MgO.
   Here's the tricky bit. You've now got to be able to write this down:

48 g of Mg ......reacts to give.....80 g of MgO
1 g of Mg ......reacts to give.....
60 g of Mg .....reacts to give......

The big clue is that in the question they've said we want to burn "60 g of magnesium", i.e. they've told us how much magnesium to have, and that's how you know to write down the left-hand side of it first, because:

You'll first need to ÷ by 48 to get 1 g of Mg
and then need to × by 60 to get 60 g of Mg.

Then you can work out the numbers on the other side (shown in orange below) by realising that you must divide both sides by 48 and then multiply both sides by 60.

÷48 48 g of Mg .... 80 g of MgO ÷48
1 g of Mg .... 1.67 g of MgO
×60 60 g of Mg .... 100 g of MgO ×60

The mass of product is called the yield of a reaction. You should realise that in practice you never get 100% of the yield, so the amount of product will be slightly less than calculated (see p63).

This finally tells us that 60 g of magnesium will produce 100 g of magnesium oxide.
If the question had said, "Find how much magnesium gives 500 g of magnesium oxide", you'd fill in the MgO side first, because that's the one you'd have the information about. Got it? Good stuff.

Reaction mass calculations — no worries, matey...

The only way to get good at these is to practise. So have a go at these: Answers on page 124

1) Find the mass of calcium which gives 30 g of calcium oxide (CaO) when burnt in air.
2) What mass of fluorine fully reacts with potassium to make 116 g of potassium fluoride (KF)?

The Mole

The mole is really confusing. I think it's the word that puts people off. It's very difficult to see the relevance of the word "mole" to different-sized piles of brightly coloured powders.

"THE MOLE" is Simply the Name Given to a Certain Number

Just like "a million" is this many: 1 000 000; or "a billion" is this many: 1 000 000 000,
so "a mole" is this many: 602 300 000 000 000 000 000 000 or 6.023×10^{23}.

1) And that's all it is. Just a number. The burning question, of course, is why is it such a silly long one like that, and with a six at the front?
2) The answer is that when you get precisely that number of atoms of carbon-12 it weighs exactly 12 g. So, get that number of atoms or molecules, of any element or compound, and conveniently, they weigh exactly the same number of grams as the relative atomic mass, A_r (or M_r) of the element (or compound). This is arranged on purpose of course, to make things easier.

One mole of atoms or molecules of any substance will have a mass in grams equal to the relative formula mass (A_r or M_r) for that substance.

Examples:

Iron has an A_r of 56.	So one mole of iron weighs exactly 56 g
Nitrogen gas, N_2, has an M_r of 28 (2×14).	So one mole of N_2 weighs exactly 28 g
Carbon dioxide, CO_2, has an M_r of 44.	So one mole of CO_2 weighs exactly 44 g

This means that 12 g of carbon, or 56 g of iron, or 28 g of N_2, or 44 g of CO_2, all contain the same number of particles, namely one mole or 6×10^{23} atoms or molecules.

Nice Easy Formula for Finding the Number of Moles in a Given Mass:

$$\text{NUMBER OF MOLES} = \frac{\text{Mass in g (of element or compound)}}{M_r \text{ (of element or compound)}}$$

Example: How many moles are there in 42 g of carbon?
Answer: No. of moles = Mass (g) / M_r = 42/12 = 3.5 moles. Easy.

"Relative Formula Mass" is Also "Molar Mass"

1) You've been quite happy using the relative formula mass, M_r, all through the calculations.
2) In fact, that was already using the idea of moles because M_r is actually the mass of one mole in g, or as we sometimes call it, the molar mass.

A "1M Solution" Contains "One Mole per Litre"

The 'moles per litre' of a solution is sometimes called its 'molarity'.

This is pretty easy. So a 2 M solution of NaOH contains 2 moles of NaOH per litre of solution.
You need to know how many moles there'll be in a given volume:

$$\text{NUMBER OF MOLES} = \text{Volume in Litres} \times \text{Moles per Litre of solution}$$

Example: How many moles in 185 cm^3 of a 2 M solution? Ans: $0.185 \times 2 = 0.37$ moles

7 moles of moles ≈ 1 Earth.....assuming vol. of 1 mole = ¼ litre, no gaps between moles, spherical Earth...

It's possible to do all the calculations on the previous pages without ever talking about moles.
You just concentrate on M_r and A_r instead — M_r and A_r represent the mass of one mole anyway.
Learn both the equations above. They'll make your life more complete (and be useful in the exam).

Percentage Yield

Percentage yield tells you about the overall success of an experiment. It compares what you think you should get (predicted yield) with what you get in practice (actual yield).

Percentage Yield Compares Actual and Predicted Yields

The more reactants you start with, the higher the actual yield will be — that's pretty obvious. But the percentage yield doesn't depend on the amount of reactants you started with — it's a percentage.

1) The predicted yield of a reaction can be calculated from the balanced reaction equation (see page 60).
2) Percentage yield is given by the formula:

$$\text{percentage yield} = \frac{\text{actual yield (grams)}}{\text{predicted yield (grams)}} \times 100$$

3) Percentage yield is always somewhere between 0 and 100%.
4) A 100% percentage yield means that you got all the product you expected to get.
5) A 0% yield means that no reactants were converted into product, i.e. no product at all was made.

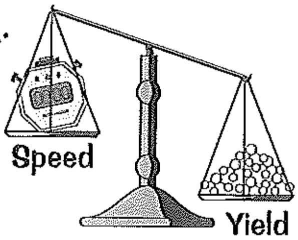
Yields are Always Less Than 100%

In real life, you never get a 100% percentage yield. Some product or reactant always gets lost along the way — and that goes for big industrial processes as well as school lab experiments. How this happens depends on what sort of reaction it is and what apparatus is being used.

Lots of things can go wrong, but the four you need to know about are:

1) The Reaction is Reversible

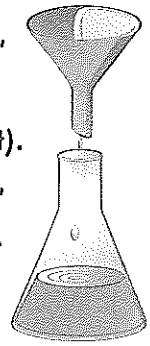
In reversible reactions (like the Haber process, see page 81), not all the reactants change into product. Instead, you get reactants and products in equilibrium. Increasing the temperature moves the equilibrium position (see page 80), so heating the reaction to speed it up might mean a lower yield.



2) Filtration

When you filter a liquid to remove solid particles, you nearly always lose a bit of liquid or a bit of solid.

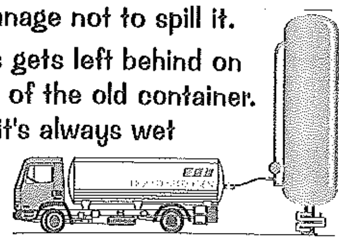
1) If you want to keep the liquid, you lose the bit that remains with the solid and filter paper (as they always stay a bit wet).
2) If you want to keep the solid, some of it usually gets left behind when you scrape it off the filter paper — even if you're really careful.



3) Transferring Liquids

You always lose a bit of liquid when you transfer it from one container to another — even if you manage not to spill it.

Some of it always gets left behind on the inside surface of the old container. Think about it — it's always wet when you finish.



4) Unexpected Reactions

Things don't always go exactly to plan.

Sometimes you get unexpected reactions happening, so the yield of the intended product goes down. These can be caused by impurities in the reactants, but sometimes just changing the reaction conditions affects what products you make.

You can't always get what you want...

A high percentage yield means there's not much waste — which is good for preserving resources, and keeping production costs down. If a reaction's going to be worth doing commercially, it generally has to have a high percentage yield or recyclable reactants, e.g. the Haber process.

Reversible Reactions

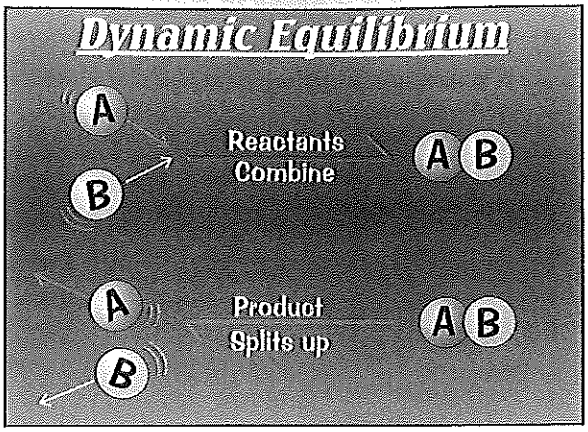
A reversible reaction is one where the products of the reaction can react with each other and convert back into the original reactants. In other words, it can go both ways.

A reversible reaction is one where the products of the reaction can themselves react to produce the original reactants

$$A + B \rightleftharpoons C + D$$

Reversible Reactions Will Reach Dynamic Equilibrium

1) If a reversible reaction takes place in a closed system then a state of equilibrium will always be reached.
2) Equilibrium means that the relative (%) quantities of reactants and products will reach a certain balance and stay there. (A 'closed system' just means that none of the reactants or products can escape.)
3) It is in fact a DYNAMIC EQUILIBRIUM, which means that the reactions are still taking place in both directions, but the overall effect is nil because the forward and reverse reactions cancel each other out. The reactions are taking place at exactly the same rate in both directions.



Changing Temperature and Pressure to Get More Product

1) In a reversible reaction the 'position of equilibrium' (the relative amounts of reactants and products) depends very strongly on the temperature and pressure surrounding the reaction.
2) If you deliberately alter the temperature and pressure you can move the 'position of equilibrium' to give more products and fewer reactants.

Temperature

All reactions are exothermic in one direction and endothermic in the other.

If you raise the temperature, the endothermic reaction will increase to use up the extra heat.

If you reduce the temperature the exothermic reaction will increase to give out more heat.

Pressure

Many reactions have a greater volume on one side, either of products or reactants (greater volume means there are more molecules and less volume means there are fewer molecules).

If you raise the pressure it will encourage the reaction which produces less volume.

If you lower the pressure it will encourage the reaction which produces more volume.

Adding a CATALYST doesn't change the equilibrium position:

1) Catalysts speed up both the forward and backward reactions by the same amount.
2) So, adding a catalyst means the reaction reaches equilibrium quicker, but you end up with the same amount of product as you would without the catalyst.

Remember — catalysts DON'T affect the equilibrium position...

Changing the temperature always changes the equilibrium position, but that's not true of pressure. If your reaction has the same number of molecules on each side of the equation, changing the pressure won't make any difference at all to the equilibrium position (it still affects the rate of reaction though).

The Haber Process

This is an important industrial process. It produces ammonia (NH_3), which is used to make fertilisers.

Nitrogen and Hydrogen are Needed to Make Ammonia

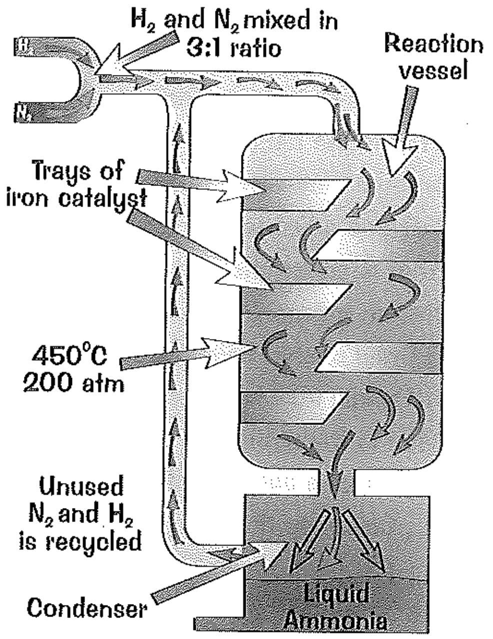
$$N_{2\,(g)} + 3H_{2\,(g)} \rightleftharpoons 2NH_{3\,(g)} \quad (+\ \text{heat})$$

1) The nitrogen is obtained easily from the air, which is 78% nitrogen (and 21% oxygen).
2) The hydrogen comes from natural gas or from other sources like crude oil.
3) Because the reaction is reversible — it occurs in both directions — not all of the nitrogen and hydrogen will convert to ammonia. The reaction reaches a dynamic equilibrium.

Industrial conditions:
Pressure: 200 atmospheres; Temperature: 450 °C; Catalyst: Iron

The Reaction is Reversible, So There's a Compromise to be Made:

1) Higher pressures favour the forward reaction (since there are four moles of gas on the left-hand side, for every two moles on the right).
2) So the pressure is set as high as possible to give the best % yield, without making the plant too expensive to build (it'd be too expensive to build a plant that'd stand pressures of over 1000 atmospheres, for example). Hence the 200 atmospheres operating pressure.
3) The forward reaction is exothermic, which means that increasing the temperature will actually move the equilibrium the wrong way — away from ammonia and towards N_2 and H_2. So the yield of ammonia would be greater at lower temperatures.
4) The trouble is, lower temperatures mean a slower rate of reaction. So what they do is increase the temperature anyway, to get a much faster rate of reaction.
5) The 450 °C is a compromise between maximum yield and speed of reaction. It's better to wait just 20 seconds for a 10% yield than to have to wait 60 seconds for a 20% yield.



6) The ammonia is formed as a gas but as it cools in the condenser it liquefies and is removed.
7) The unused hydrogen, H_2, and nitrogen, N_2, are recycled so nothing is wasted.

The Iron Catalyst Speeds Up the Reaction and Keeps Costs Down

1) The iron catalyst makes the reaction go faster, which gets it to the equilibrium proportions more quickly. But remember, the catalyst doesn't affect the position of equilibrium (i.e. the % yield).
2) Without the catalyst the temperature would have to be raised even further to get a quick enough reaction, and that would reduce the % yield even further. So the catalyst is very important.

You need to learn this stuff — go on, Haber go at it...

The trickiest bit is remembering that the temperature is raised not for a better equilibrium, but for speed. It doesn't matter that the percentage yield is low, because the hydrogen and nitrogen are recycled. Cover the page and scribble down as much as you can remember, then check, and try again.