NASCA WORKBOOK

Chemistry Section 2

Introduction/Overview

Chemical changes happen within living organisms and in non-living systems. The change from reactants (starting materials) to products can be described using chemical equations involving words, models or chemical symbols. Learning the symbolic language need to describe chemical reactions is essential for any chemist.

Due to the laws of constant composition and conservation of mass, chemical equations must be balanced by means of a chemical accounting system. This allows chemists to do chemical calculations. For example, how much of a starting material do I need to make a certain amount of product with minimum wastage? As you can imagine, this is an important question in industry in order to keep costs down. You will learn how to do this accounting, calculating quantities such as such as molar concentrations, limiting reagents and percentage yield.

Chemical changes involve the breaking of existing chemical bonds (with an input of energy) and a release of energy when new chemical bonds are formed. Depending on the relative amounts of energy absorbed or energy released, a chemical reaction can be described as exothermic or endothermic. The changes in potential energy over the course of chemical reactions can be represented using potential energy profiles. It is important to develop an understanding of these energy changes that take place during chemical reactions.

Some chemical reactions (such as those in a fireworks display) happen very quickly while others may take place very slowly (such as the rusting or iron). The rate of a chemical reaction is the change in reactant/product concentration per unit time and this can be measured in various ways depending on the characteristics of a particular reaction. In this section, the factors that influence the rate of a chemical reaction will be explored. Again, this is important when you want to produce a product for sale, or prevent a reaction from happening for as long as possible (such as rusting of your gate).

In a closed system, some chemical reactions are reversible. This means that reactants can go to products, but products can also break down and return to the reactant form. If such reactions are left undisturbed, a dynamic equilibrium will be reached in which the rate of the forward and the reverse reactions are equal. Manipulating the conditions of a reaction enables chemists to determine the results of that reaction. Again, this has important industrial applications in order to maximize product formation and therefore profit.

There are many different types of chemical reactions such as substitution reactions, synthesis, redox reactions etc. In this course there is not enough time to look at these, but you will explore acids and bases and neutralization reactions.

Subtopic 2.1 Chemical change

At the end of this subtopic you should be able to:

2.1.1 Represent chemical reactions using words and chemical symbols (including state symbols);
2.1.2 State the Laws of Constant Composition and Conservation of Matter;
2.1.3 Balance chemical equations correctly.

Whenever new chemical substances are formed a chemical change has taken place.

If your eyes were strong enough to actually see what happens during a chemical change you would see that bonds in the reactants are broken and new bonds form to create the new products.

Reactants: these are the particles (elements and compounds) present at the start of the reaction. So they are the starting materials of a chemical reaction.

Products: these are the particles (elements and compounds) produced as a result of the chemical reaction.

Some examples of chemical changes that you are already aware of are as follows:

- The rusting of iron
- The burning of fuels
- Digestion of food
- Cooking food
- Rotting matter

In order to describe and work with chemical reactions you need a way to write them. This requires an understanding of several things:

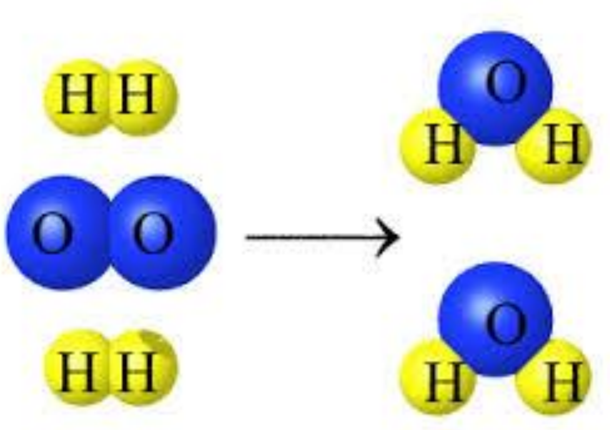
- An understanding of the Law of Constant Composition.
- An understanding of the Law of Conservation of Matter.
- The chemical formulae of all the reactants and all the products.
- The physical state in which the reactants and products exist.
- The amounts and ratios of reactants required for a complete reaction to take place and the amounts and ratios of the products formed.

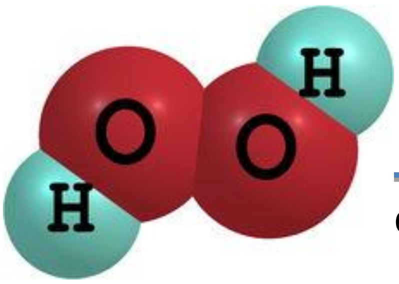
Unit 2.1.1 Law of Constant Composition

The law: All samples of a particular chemical compound have the same elemental composition by mass regardless of how the compound was made or where it is found.

For example, water is always made from 2 atoms of hydrogen (H) bound to 1 atom of oxygen (O) and the mass ratio of H to O is always 11%:89%.

This is always true for water. If the proportions change the substance has changed and is no longer water.





It is possible for H and O to combine in a different ratio, namely H_2O_2. This is no longer water, but a different compound called hydrogen peroxide.

Why is the Law of Constant Composition useful?

If you can measure the mass of each element present in a compound, you can determine the chemical formula of that compound.

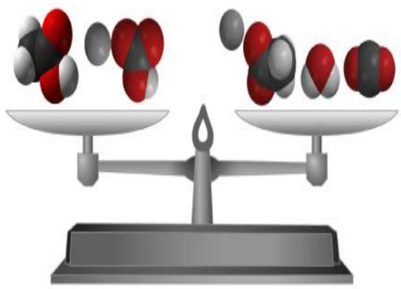
Activity 2.1.1: Answer the following questions

1. In carbon dioxide (CO_2) the ratio of mass of C:O is always 3:8. If 6.0 g of carbon is completely converted into carbon dioxide, how much oxygen (in grams) is required?
2. The mass ratio of H:O in water is 1:8. In a sample of water that has a mass of 36 g what is the mass of H and O?

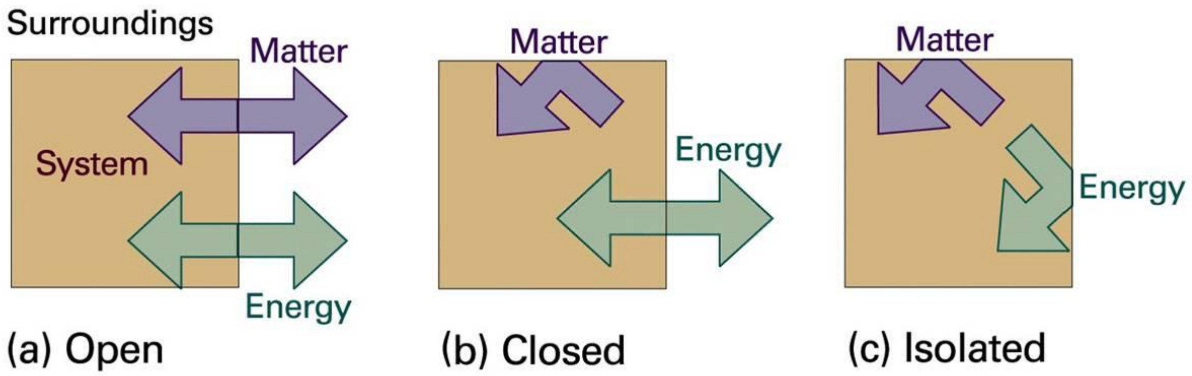
Unit 2.1.2 Law of Conservation of Matter

The law: Mass, in a closed system, is neither created nor destroyed by chemical reactions or physical changes. Therefore, according to the law of conservation of mass, the mass of the products in a chemical reaction must equal the mass of the reactants.

A simpler way of saying this is that mass is never lost or gained during a chemical reaction, it simply changes its form (form reactants to products). The total mass of reactants is equal to the total mass of products. Or, the total mass of starting materials is equal to the total mass of substances after the reaction is complete.



Note: If one of the products is a gas these measurements have to be made in a closed system to make sure that no reactants or products escape from the reaction vessel. If that happened the masses would appear to be different. In an open system, both energy and matter can escape from the container to the surroundings. In a closed system, energy can still escape but matter e.g. a gas, remains within the container. There is a third type of system, the isolated system which keeps both energy and matter from escaping into the surroundings.



Why is the Law of Conservation of Matter useful?

It means that if you can measure the masses of the reactants before a reaction starts and the mass of one of the products, you can calculate the mass of the other product without having to measure it.

It also means that you can measure the information you need to write down chemical equations that describe exactly what the reactants and products in a reaction are, what the mass ratios are.

Activity 2.1.2: Answer the following questions

1. In a chemical reaction, 300 grams of reactant A are combined with 100 grams of reactant B. Both A and B react to completion so no A or B is left over. How much will the total product weigh?
2. In a reaction, 25 grams of reactant AB breaks down completely into 10 grams of product A and an unknown amount of product B. How much does product B weigh?
3. Consider the following reaction:

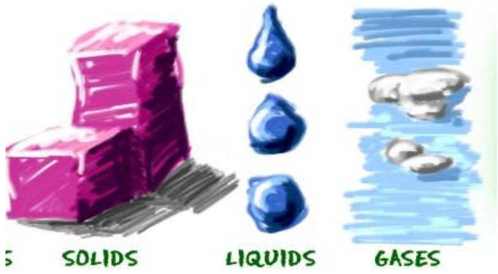
   $$CaCl_2 + Na_2CO_3 \rightarrow NaCl + CaCO_3$$

   If 2.12 g of $CaCl_2$ reacts completely with 2.20 g of Na_2CO_3 and

2.90 g of calcium carbonate is formed, what amount of NaCl is produced?
4. In an experiment you react 40.0 g of calcium with 16.0 g of oxygen. At the end of the reaction you have calcium oxide. What would the mass of the product be?
5. If 24.0 g of magnesium reacts with oxygen to form 40.0 g of magnesium oxide, how much oxygen was used up in the reaction?

2.1.3 State symbols

State symbols allow you to show the physical state in which the reactants and products are present during a chemical reaction.



The state symbols are as follows:

- **g** for gas
- **s** for solid
- **l** for liquid
- **aq** for aqueous solution (a solution in water)

State symbols are shown in brackets after the chemical formula.
For example:

- Water, a liquid, is shown as $H_2O(l)$
- Copper, a solid is shown as Cu(s)
- Nitrogen, a gas, is shown as $N_2(g)$
- A solution of sodium chloride in water is shown as NaCl(aq)

Unit 2.1.4 Representing chemical reactions

Chemical reactions are represented by chemical equations.

- **Chemical equations can be word equations**. For example:

Carbon reacts with oxygen to form carbon dioxide

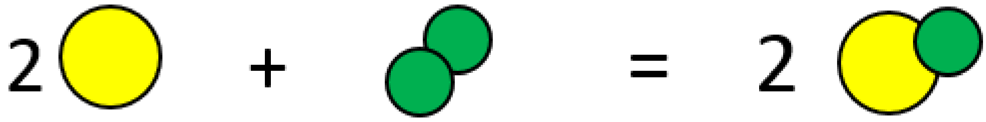
- In chemistry, symbol equations are usually preferred as they contain more useful information. In a **symbol equation** the word equation given above becomes:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

As you can see there is now information about the atoms making up the reactants and the products, the ratios in which they combine and the state in which each component is when the reaction takes place. You will also learn more about the importance of equations of this type as you go along.

Unit 2.1.5 Balancing chemical equations

The total number of atoms of each element present in the reactants must always be equal to the total number of the same elements in the products, regardless of how they are arranged into compounds. This is to obey the Law of Conservation of Mass. No atoms can disappear during the reaction, nor can additional atoms be produced.



Consider the chemical equation for the reaction that forms water.

To form water you know that you need hydrogen gas and oxygen gas to react together to form the water. You also know the formula of these molecules: $H_2(g)$, $O_2(g)$ and $H_2O(l)$.

The chemical equation for the formation of water is therefore:

$$H_2(g) + O_2(g) \rightarrow H_2O(l)$$

But wait a minute, does this obey the Law of Conservation of Mass?

On the left hand side there are 2 H atoms and 2 O atoms.

On the right hand side there are 2H atoms but only 1 O atom, so something appears to have been lost during the reaction. This cannot be according to the law.

So the equation is *not balanced* and you have to do something about that.

A water molecule has two H atoms for every O atom, so the ratio is 2:1 for H:O. Therefore, on the left hand side the ration also has to be 2:1.

Now the chemical equation looks like this:

$2H_2(g) + O_2(g)$ …… $H_2O(l)$

Let's check if this now works:

On the left: 4 H atoms and 2 O atoms

On the right: 2H atoms and 1 O atom.

Still wrong.

To balance the equation correctly the number of atoms of both H and O on the right must be doubled.

Now the chemical equation looks like this:

$2H_2(g) + O_2(g)$ …… $2H_2O(l)$

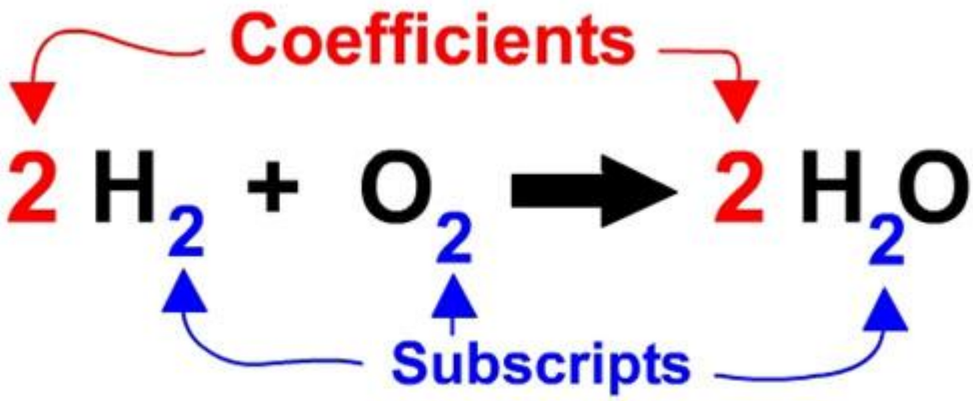
Let's check if this now works:

On the left: 4 H atoms and 2 O atoms

On the right: 4H atoms and 2 O atoms.

We have a balanced equation for the reaction!

In order to balance equations in this way remember that only the numbers in front of the formula can be changed. These are called **stoichiometric coefficients***.* The subscripts must stay the same.



Note: If you change the numbers after the element symbols (the subscripts) you will be changing the nature of the substance and this is not allowed.

Activity 2.1.3: Write *word equations* and *balance* the following equations:

(Remember that ions such as NO_3^-, OH^-, HCO_3^- and CO_3^- move as complete units).

1. $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
2. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
3. $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
4. $NH_3(g) \rightarrow N_2(g) + H_2(g)$
5. $NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$

Subtopic 2.2 Quantitative aspects of chemical change (Stoichiometry)

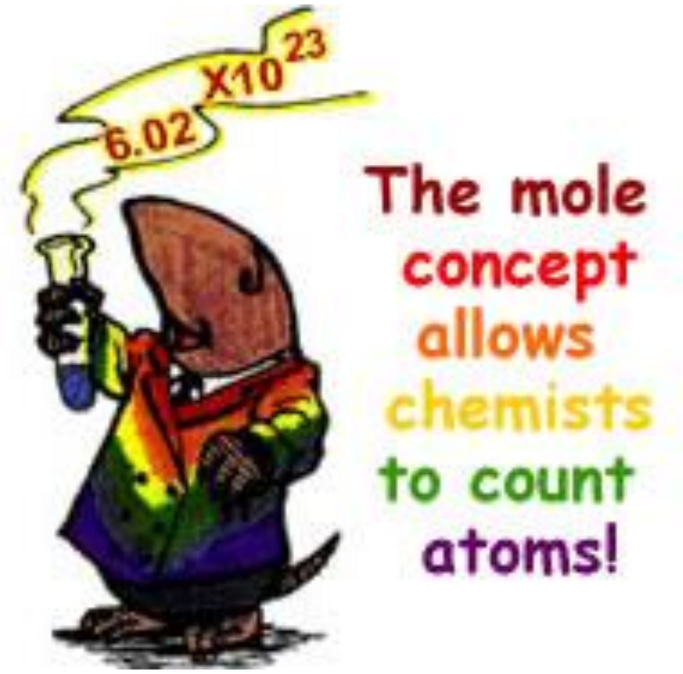
At the end of this subtopic you should be able to:

2.2.1 Calculate number of moles, volume, concentration, limiting reagent and percentage yield for chemical reactions;
2.2.2 Determine the empirical and molecular formulae when given the percentage composition of a particular compound.
2.2.3 Define limiting reagent, theoretical yield, actual yield and percentage yield;

Quantitative means amounts or quantities. In chemistry we might need to ask the following questions:

- Can we count atoms?
- How could we measure known amounts of atoms without counting?
- How could we measure the amounts o compounds?
- What does a formula tell us about the make-up of a compound?
- What do chemical equations tell us about the amounts of reactants and products involved in chemical reactions?

Unit 2.2.1 Relating number and mass



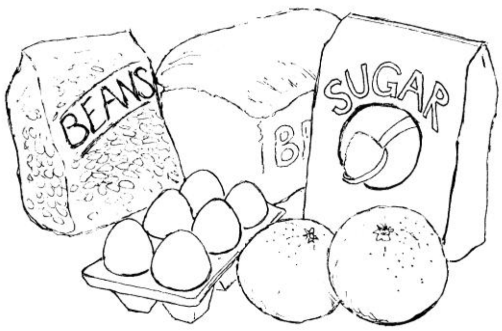
You have learned that atoms are very small – they have an approximate diameter of about 0.000 000 000 01 metres and an average mass of about 0.000 000 000 000 000 000 000 01 g! It is this very small mass in grams that is called 1 atomic mass unit (amu). So there really is a big problem facing us when we want to know how many atoms are involved in the formation of compounds and in chemical reactions because it is impossible to count individual atoms or weigh them.

Can you suggest a possible way in which this problem can be solved?

Did you know: If you take about 100 seconds to count 100 atoms it would take you

95 000 000 000 000 000 years (9.5×10^{17}) to count the atoms in a 500g packet, counting 24 hours a day and 7 days a week! You wouldn't even have time to have a drink!

Activity 2.2 1: Counting small things



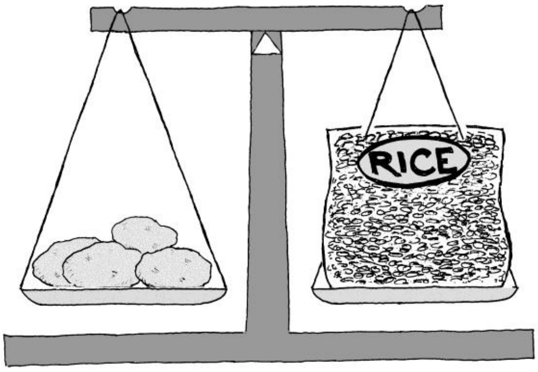
A variety of groceries such as a packet of rice, a packet of beans, a bag of apples, a tray of eggs, a packet of granulated sugar was collected. From each product 12 units were counted out and weighed e.g. 12 grains of rice. The results are recorded in the table below.

(Note: Scientific balances were used to weigh small amounts and small masses).

Product	Mass of 12 units (g)	Average mass of 1 unit (g)	Mass of the bag of product (g)	Number of units per bag	Number of units in a 1000 g bag of product
e.g. Rice	(e.g. 12 grains of rice)	(e.g. 1 grain of rice)	(e.g. mass of whole bag of rice)	(e.g. number of grains of rice in the bag)	(e.g. number of grains of rice in a bag weighing 1000 g)
Rice	0.15				
Sugar	0.02				
Apples	1417.80				
Dried beans	11.64				
Eggs	735.36				

1. What is the word we use to describe 12 of anything?
2. Work out (by calculation, not by weighing), the average mass of a single unit of each product e.g. a single grain of rice, and add this information to your table.
3. Now calculate the number of units in each bag, and add this information to your table.
4. Now calculate the number of units in a 1000 g bag of each product and add the information to the table.
5. Now answer the following questions:
    a. Why does one use mass to measure out a product rather than number?
    b. Why is it reasonable to buy items like eggs or apples by number rather than mass whereas we buy goods like sugar and rice by mass, not number?
    c. How does the number of units vary for packets of 1000 g of different items? Why?
    d. To be sure that you have the same number of items of different products would you weigh the same masses of each item, or different masses?
    e. If you wanted to supply different customers with the same numbers of units of a single product would you measure the same mass for each customer or different masses?

f. If one atom has an average mass of 0.000 000 000 000 000 000 000 01 g (1×10^{-23} g) how many atoms must one weigh to achieve a mass of 1.00 g?
g. Why is it impossible to count atoms?



Got it: If particles are small it is easier to measure them by mass than by counting them. This is true of regular grocery items like rice and beans as well as much smaller particles like atoms. Approximately 100 000 000 000 000 000 000 000 (1×10^{23}) atoms would be contained in 1 g of atoms, depending on the atomic mass of the atoms in question.

Activity 2.2.2: Relating atomic number and mass

1. Study at the table below and think about the questions that follow.

Number of atoms of each element present	Mass of carbon atoms	Mass of hydrogen atoms	Mass of oxygen atoms	Mass of sodium atoms
1	12.0 amu	1.0 amu	16.0 amu	23.0 amu
2	24.0 amu	2.0 amu	32.0 amu	46.0 amu
50	600.0 amu	50.0 amu	900.0 amu	1150.0 amu
1000	12000.0 amu	1000.0 amu	16000.0 amu	23000.0 amu
6.02×10^{23}	12.0 g	1.0 g	16.0 g	23.0 g
1.20×10^{24}	24.0 g	2.0 g	32.0 g	46.0 g
6.02×10^{24}	120.0 g	10.0 g	160.0 g	230.0 g

Note: amu = atomic mass unit = the unit of mass for the very small atom, not measurable with our technology = 1×10^{-23} g approximately.

a. What do you notice about the mass of each element and the atomic mass of that element on the Periodic Table?
b. What do you notice about the mass in grams of 6.02×10^{23} atoms of a particular element and the mass of one atom of that element?
c. Can you make a general statement about the mass of 6.02×10^{23} atoms of a particular element?
d. Why is it useful to have this number of 6.02×10^{23}?

Did you know: The number 6.02×10^{23} which is also called *Avogadro's number*, has a special name – the **mole.** This name is the same sort of name as is a dozen for the number 12. Can you think of other names for particular numbers?

Got it: One mole of an element always contains 6.02×10^{23} atoms of that element. It has a mass in grams that has the same numerical value as the relative mass of one atom (in amu).

1 mole = 6.02×10^{23} particles = Avogadro's number of

Activity 2.2.3: Test your knowledge of mass, moles and Avogadro's Number

1. What is the mass of 1 mole of calcium atoms?
2. What is the mass of 0.235 moles of nickel atoms?
3. How many atoms would you find in 0.250 moles of chlorine atoms?
4. How many atoms would you find in 3.2 moles of arsenic atoms?
5. What is the mass of 3.02×10^{23} atoms of aluminium?

6. What is the mass of 10 000 atoms of caesium?

Remember: The mass of 1 mole of an element is called its **relative molar mass**

Insert comical pic of mole weighing things – balanced scale, but sizes clearly different

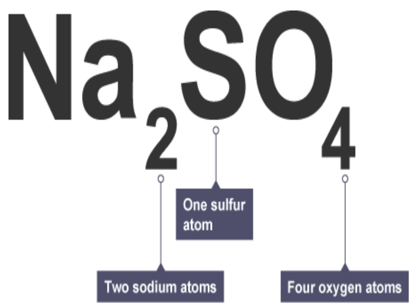
Got it: 1 mole = 6.02×10^{23} particles and 1 mole of an atom has a mass equal to its relative atomic mass expressed in grams.

$$\text{No of moles} = \frac{\text{actual number of particles present}}{6.02 \times 10^{23} \text{ particles}}$$

Ask yourself: Can you determine the number of particles present in any number of moles of any substance?

Unit 2.2.2 Reading formulae

A formula is a shorthand symbol of something. In order to understand the information carried in a formula we need to know certain things. For example, what does the formula CO_2 tell us? What about the formulae NaOH and H_2SO_4?



A formula tells us the *combining ratio* of the elements that make up a compound. Therefore we see that carbon dioxide is made up of carbon and oxygen in a ratio of 1:2. This ratio can be seen as one C atom for every two O atoms, or one dozen C atoms for every two dozen O atoms, or one mole of C atoms for every two moles of O atoms. Can you explain this statement?

NaOH tells us that sodium hydroxide formula units are made up of 1 sodium atom, 1 oxygen atom and 1 hydrogen atom. What would be the mole ratio of these elements in NaOH?

H_2SO_4 tells us that sulfuric acid is made up of 2 hydrogen atoms, 1 sulfur atom and 4 oxygen atoms. What would be the mole ratio of the elements in H_2SO_4?

Na_2SO_4 tells us that sodium sulphate is made up of 2 sodium atoms, 1 sulfur atoms and 4 oxygen atoms. What would be the mole ratio of the elements in Na_2SO_4?

Activity 2.2 4: If you know the formula of a compound, would it be possible to calculate the molar mass of the compound? What about the masses of each of the contributing elements? Use the three examples above (CO_2, NaOH and H_2SO_4) to answer the question.

Got it: The mass of one mole of any compound is equal to the sum of the masses of the constituent elements The masses of the elements are obtained by counting the number of atoms of that element present and multiplying the relative atomic mass value by that number. This is summarised in the following example:

For one mole of calcium carbonate ($CaCO_3$)				
Number of atoms per formula unit	Number of moles of atoms	No of atoms in 1 mole of compound	RAM (amu)	Mass of atoms (g/mol)
Ca: 1	1	6.02×10^{23}	40.08	40.08 (1 x 40.08)
C: 1	1	6.02×10^{23}	12.01	12.01 (1 x 12.01)
O: 3	3	1.81×10^{24}	16.00	48.00 (3 x 16.00)
No of atoms in 1 mole		3.01×10^{24}		
Molar mass of $CaCO_3$				100.09

number of moles = $\frac{\text{actual mass of sample}}{\text{RMM of sample}}$

Ask yourselves: What would be the difference between a molecular weight and a formula weight?

Answer: Molecules are the units of compounds made up from non-metals combing with each other while formula units indicate compounds made up from metals and non-metals. Therefore a molecular weight is the molar mass of a molecule and a formula weight is the molar mass of a formula unit. CO_2 and H_2SO_4 would have molecular weights while NaOH would have a formula weight. Actually most chemists have become careless and use the term molecular weight most of the time.

Activity 2.2.5: Working with formula and molar masses

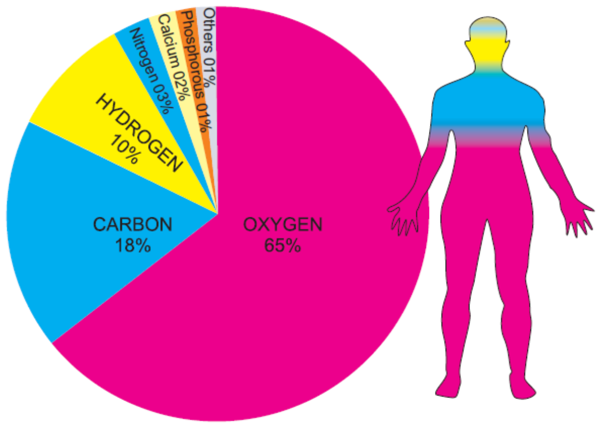
1. Answer the following questions and place your answers in a table like the one suggested below:

Compound formula	Formula unit or molecule?	Molar mass of compound (g)	Atomic mass of first element in the formula	Atomic mass of second element in the formula	% (by mass) of first element	% (by mass) of the second element	Total %	Mass of 1st element in 100 g sample	Mass of 2nd element in 100 g sample
CO									
$AlCl_3$									
H_2O									
Fe_2O_3									

   a. Calculate the molar masses of each of the compounds listed
   b. Note whether the formula represents a formula unit or a molecule
   c. Note the contributing mass of each element in the compound towards the molar mass
   d. Calculate the percentage mass of each contributing element relative to the molar mass
   e. Calculate the sum of the two percentages for each formula and insert into table. What do you notice?

f. Using CO as your example explain what is meant by % (by mass) of C and O
g. Do you think that this percentage will stay constant for CO? What about the percentages for the other examples?
h. Formulate a general statement concerning the percentage composition (by mass) of elements in a compound
i. Calculate the mass of each element in a 100 g sample of the compound in which it occurs
j. If a compound was composed of 4 different elements how many % composition values would describe the composition of the compound?

Got it: **Percentage composition** is the mass of each element per 100 mass units of the compound in which they are found. The percentage composition is *constant* for a particular compound and can therefore be used to determine the mass of each element in any given mass of the compound.



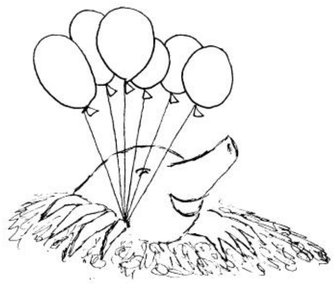
$$\% \text{ composition of element by mass} = \frac{\text{RAM (g) of element}}{\text{RMM (g) of compound}} \times 100\%$$

Activity 2.2.6: Test your knowledge of percentage composition (by mass)

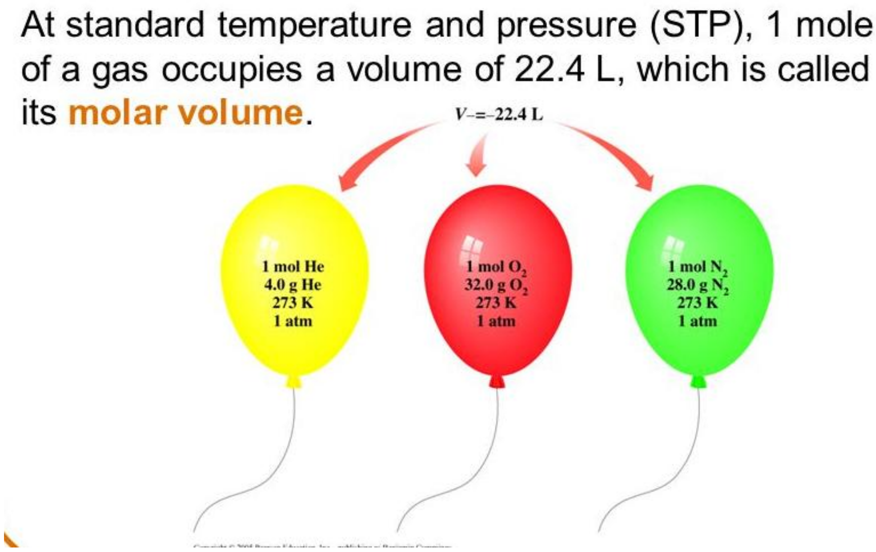
1. What mass of carbon would be present in 150.0 g of the following compounds:
   a. Carbon dioxide, CO_2
   b. Carbon monoxide, CO
   c. Methane, CH_4
   d. Calcium carbonate, $CaCO_3$
   e. Acetic acid, CH_3COOH

f. Hydrochloric acid, HCl

Unit 2.2.3 Moles and volumes of gas



In the previous section we have seen the relationship between numbers of particles, moles and mass. What about substances that occur as gases at room temperature? Measuring the mass of a gas is possible, but measuring a volume would often be much more convenient. Can we relate numbers of particles, moles and volumes?



As it happens it is possible to do so. Two gases containing the same number of particles will occupy the same volume when the conditions of temperature and pressure are the same.

This is known as **Avogadro's Theory**. From this follows the statement that 1 mole of any gas occupies the same volume as one mole of any other gas at the same temperature and pressure.

The standard temperature and pressure (STP) that are usually used are:

0^oC and 1 atmosphere of pressure.

Therefore, 1 mole of any gas occupies 22.41dm^3 at 0^{o}C and 1 atm pressure.

The equation relating number of moles to a volume of gas is as follows:

$$\text{number of moles} = \frac{\text{actual volume of sample (at STP)}}{22.41\text{dm}^3 \text{ (at STP)}}$$

Got it: If the conditions of temperature and pressure are fixed, the volume of 1 mole of any gas will be a constant number called the **molar volume**. The fixed conditions are called **Standard temperature and pressure, (STP),** which means a **temperature of 0^{o}C and a pressure of 1 atmosphere**. If one chooses different conditions e.g. a temperature of 25^{o}C, the volume of 1 mole of gas will be 24.47 dm^3, and the conditions need to be stated.

1 mole of any gas occupies a volume of 22.41 dm^3 at STP
1 mole of any gas occupies a volume of 24.47 dm^3 at 25^{o}C and 1 atm pressure

Challenge: Using the information you have just discovered, can you make a relationship between the volume of a gas (at STP, say), the number of particles of the gas and the mass of the gas?

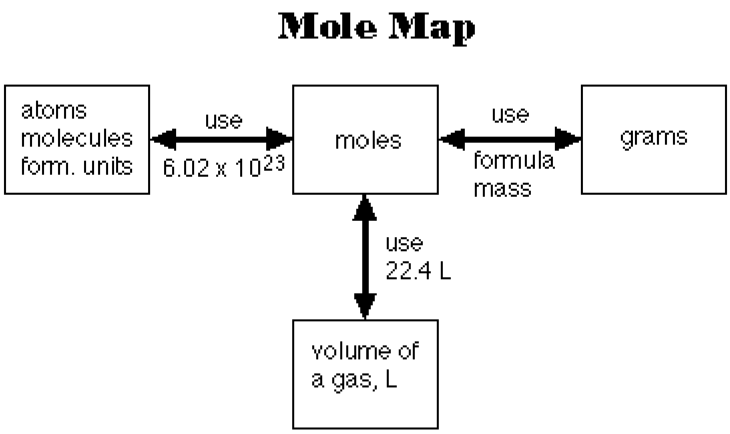
Answer:
At STP: 1 mole of gas occupies 22.41 dm^3 and contains 6.02 x 10^{23} particles with a total mass equal to the molar mass of the gas

$$\begin{array}{ccccc} 6.02 \times 10^{23} \text{ particles} & \leftarrow & 1 \text{ mole} & \rightarrow & 22.41 \text{ dm}^3 \\ & & \downarrow & & \\ & & \text{molar mass (g)} & & \end{array}$$

Activity 2.2.7: Test your understanding of the relationships between mass, moles and volume of gases.

1. What is the volume of 21.0 g of CO_2 measured at STP?
2. What is the mass of 0.0573 dm^3 of NO gas measured at 25^oC and 1 atm pressure?
3. What is the volume of 2.44 x 10^{19} molecules of hydrogen gas (H_2) at STP?
4. What volume of oxygen gas (O_2) must one measure to get 55.0 g of the gas at STP?
5. What volume of nitrogen gas (N_2) must one measure so as to have 5.0 x 10^{22} molecules of gas present at 25^oC and 1 atm pressure?

Got it: The number of particles is related to mass and volume (in the case of a gas) by means of the number of moles present.



Unit 2.2.4 Moles and volumes of aqueous solutions

Cartoon of more concentrated vs less concentrated drink?

What is meant by an aqueous solution? What is meant by the concentration of a solution? How does a concentrated solution differ from a dilute solution? An aqueous solution is one in which a solute (solid or liquid) is dissolved in water to give a homogeneous mixture. The concentration of a solution is a measure of the amount of solute (mass of solid or volume of liquid) dissolved per unit volume of water. This would be expressed as g/ml or ml/ml, for example. A concentrated solution has a greater amount of solute present per unit volume than is present in a dilute solution.

You have already seen how you can express amounts of substances in terms of moles, and that this is important because the number of moles is related to the actual number of particles present.

Ask yourselves: If the concentration of a solution can be expressed in the units of g/ml how can it also be expressed in moles/ml? Use the example of a solution of saline (NaCl in water) at a concentration of 3.0 g/ml to work with. One frequently expresses the concentration in units of mol/dm^3. Convert your answer to this unit

Answer: RMM of NaCl = 22.99 + 35.45 = 58.44 g/mol

Moles NaCl = mass/RMM = 3.0 g/58.44 g/mol = 0.051

Concentration = 0.051 mol/ml = 0.051 x 10^3 mol/1 x 10^3 ml =51 mol/dm^3

Got it: The concentration of a solution is usually taken to be the number of moles of solute per cubic decimetre of solution i.e. mol.dm$^{!3}$ or mol/dm^3.

Another name for concentration expressed in mol.dm$^{!3}$ is **molar** (M) so a 6.0 M HCl solution would be an HCl solution with a concentration of 6.0 mol.dm$^{!3}$.

$$\text{Concentration} = \frac{\text{moles of solute}}{\text{volume of solution (dm}^3\text{)}}$$

$$\text{Number of moles} = \frac{\text{mass of solute (g)}}{\text{molar mass of solute (g.mol}^{-1}\text{)}}$$

$$\text{Concentration} = \frac{\text{mass of solute (g)}}{\text{RMM of solute (g.mol}^{-1}\text{)}} \times \frac{1}{\text{vol solution (dm}^3\text{)}}$$

Activity 2.2.8: Test yourselves

1. 17.0 g of NaOH was dissolved in 0.250 dm^3 of water. Calculate the molar concentration of the solution.
2. Calculate the molar concentration (in mol.dm$^{!3}$) of a solution made by mixing 300.0 g of $CuSO_2.5H_2O$ in 750 ml of water.
3. What is the volume of a 6.0 M solution of HCl that contains 8.3 g of HCl?
4. What is the mass of NaCl present in 50.0 ml of a 3.2 mol.dm$^{!3}$ solution of sodium chloride?

5. What is the molar concentration of a solution made by dissolving 4.55 g $CaCl_2$ in 125 ml of water?

6. What mass of $FeCl_3$ is required to make 2.00 dm^3 of a 0.200 M solution?

Unit 2.2.5 Empirical and molecular formulae

An **empirical formula** gives the **smallest whole-number ratio** of atoms in a compound. For example, the formula for butane is C_4H_{10} and the empirical formula for butane is C_2H_5.

A **molecular formula** gives the **actual number of atoms** of each of the elements present in a molecule of a specific compound. The molecular formula for butane is C_4H_{10}.

Activity 2.2.9: What is the empirical formula of the following compounds?

1. C_2H_4
2. $C_{11}H_{22}O_{11}$
3. H_2O
4. $C_{25}H_{50}$

Determining an Empirical and Molecular Formula

Before you can work out the empirical formula of a compound you need to know the mass of each of the elements in 100 g of the compound.

This gives the % mass. Then there are three calculation steps that you need to do, as follows:

1. Convert the mass of each element to moles using the molar mass from the periodic table.
2. Divide each mole value by the smallest number of moles calculated.
3. Round to the nearest whole number.

Example 1:

In an experiment a compound is found to contain 50.05 % sulfur and 49.95 % oxygen by weight. What is the empirical formula for this compound? The molecular weight (molar mass) for this compound is 64.07 g/mole

Elements	S	O
% mass	50.05	49.95
Mass (g) in 100g sample	50.05/100 x 100g = 50.05	49.95/100 x 100g = 49.95
Molar mass (g/mole)	32.07	16.01
No of moles = mass/molar mass	50.05/32.07 = 1.5608	49.95/16.00 = 3.1212
Divide both sides by lower value to give a whole number of 1 for one element at least	1.5608/1.5608 = 1	3.1212/1.5608 = 2
Since these are both whole numbers and the ratio cannot be reduced to smaller whole numbers the empirical formula can be written now	SO_2	

Determining the molecular formula:

Remember that the empirical formula is the lowest whole number ratio of elements in the compound while the molecular formula is the actual numbers of component elements present.

To determine whether the molecular formula is the same as or different from the empirical formula the following calculation needs to be done:

1. Determine the molar mass of the empirical formula, called the empirical formula mass (EFM):

   EFM of SO_2 = 32.07 + 16.01 + 16.01 = 64.09 g/mole

2. Divide the measured molecular weight by the empirical formula weight:

   64.07/64.09 = 1

3. The EFM and the molar mass are the same, so the empirical formula and the true molecular formula are the same: SO_2

Example 2:

A compound is found to contain 48.38% carbon, 8.12% hydrogen, and 53.5% oxygen by mass. Determine the empirical formula.

Elements	C	H	O
% mass	48.38	8.12	53.5
Mass (g) in 100g sample	48.38/100 x 100g = 48.38	8.12/100 x 100g = 8.12	53.5/100 x 100g = 53.5
Molar mass (g/mole)	12.01	1.01	16.00
No of moles = mass/molar mass	48.38/12.01 = 4.028	8.12/1.008 = 8.056	53.5/16.00 = 3.336
Divide both sides by lower value to give a whole number of 1 for one element at least	4.028/3.336 = 1.2	8.056/3.336 = 2.4	3.336/3.336 = 1
To convert all these numbers to whole numbers each one needs to be multiplied by 5	1.2 x 5 = 6	2.4 x 5 = 12	1 x 5 = 5

Empirical formula	$C_6H_{12}O_5$

Example 3: You have determined that the empirical formula of a compound is $C_{10}H_7O_2$. The compound has a molar mass of 318.31 g/mol. What is the molecular formula of this compound?

1. Determine the empirical formula mass (EFM) mass of the empirical unit.

   EFM = 10(12.00) + 7(1.008) + 2(16.00) = 159.06 g/mol

2. Determine the number of EFM amounts in the molar mass:

   Molar mass/ EPM = 318.31 g/mol / 159.06 g/mol = *2.001*

3. Molecular formula.

   Since there are two empirical units in a molecular unit, the molecular formula is:

$$C_{20}H_{14}O_4$$

Activity 2.2.10: Try the following problems for yourself.

1. A compound is found to contain 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. Its molar mass is 60.0 g/mol. What is the molecular formula of the compound?
2. A hydrocarbon is found to contain 85.7% carbon. Its molar mass is 84.0 g/mol. What is its molecular formula?
3. A compound present in iron ore is made up of 72.3% iron and 27.7% oxygen. The molar mass of this compound is 231.4 g/mol. What is its molecular formula?
4. A compound containing 40.0% carbon, 5.7% hydrogen and 53.3% oxygen has a molar mass of 175 g/mol. What is the molecular formula?
5. A compound is contains 87.4% nitrogen and 12.6% hydrogen. If the molecular mass of the compound is 32.05 g/mol, what is the molecular formula?

https://www.khanacademy.org/science/chemistry/chemical-reactions-stoichiome/empirical-molecular-formula/v/empirical-molecular-and-structural-formulas

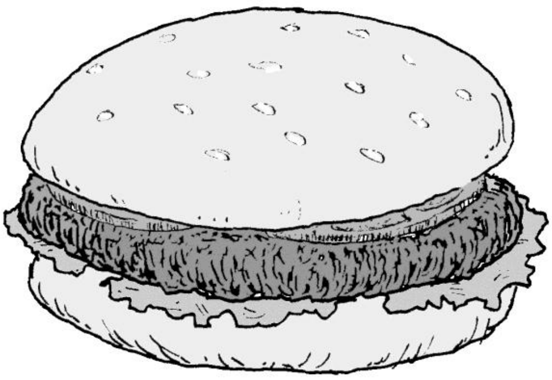
tps://www.youtube.com/watch?v=ZiXtpuDZlP8

https://www.youtube.com/watch?v=lyWAGMEKzSY

Subtopic 2.2.6 Understanding stoichiometry

Stoichiometry (pronounced stow-key-om-a-tree) is a big name (derived from the Greek) that means a kind of chemical accounting. The best way to begin to understand how this works, and why it is important, is to work with an analogy that is familiar to us.

Activity 2.2.11: What's in a hamburger? And what about the juice?



1. For the purposes of this activity assume that a hamburger needs 1 hamburger bun, 1 meat patty, 3 lettuce leaves and 2 slices of tomato. Also, 4 people will be eating together.
2. Answer the following questions regarding the hamburger:
   a. Write an equation describing the "building" of a hamburger (i.e. write the ingredients on the left hand side of the equation and the product (or result) on the right hand side in the form A + B + C + D = E
   b. If you had one hamburger for each person in the group
      i. How many hamburger buns would there be?

    ii. How many meat patties would there be?
    iii. How many lettuce leaves would there be?
    iv. How many tomato slices would there be?

c. How many hamburgers could you make if you had:
    i. 4 hamburger buns, 3 meat patties, 12 lettuce leaves and 8 slices of tomato?
    ii. 4 hamburger buns, 4 meat patties, 12 lettuce leaves and 8 slices of tomato?
    iii. 3 hamburger buns, 3 meat patties, 12 lettuce leaves and 8 slices of tomato?
    iv. 4 hamburger buns, 4 meat patties, 10 lettuce leaves and 7 slices of tomato?
    v. 4 hamburger buns, 4 meat patties, 11 lettuce leaves and 9 slices of tomato?

d. If each meat patty is made from 200 g of beef mince, and you are provided with 1,500 kg of mince, how many hamburgers could you make? Assume that all the other ingredients are freely available.

e. If you are provided with 10 meat patties but burn three of them while you are cooking them what percentage of hamburgers will you be able to make?

3. What have you found regarding the proportions of each ingredient for making a whole hamburger?
4. What have you found in the relationship between mass (of mince) and number of patties (and therefore hamburgers) that can be produced? How does this relate to the relationship between the mole and relative molar mass and Avogadro's number?
5. What is the wastage in your hamburger preparation through careless burning of some of the patties? How does this relate to the concept of percentage yield? Would an employer at a hamburger shop be pleased with your performance?
6. Now look at the simple equation that follows and answer the questions:

$$H_2(g) \quad + \quad O_2(g) \quad \rightarrow \quad H_2O(l)$$

    a. What does the formula H_2 say about the gas hydrogen?
    b. What does the formula O_2 say about the gas oxygen?
    c. What does the formula H_2O say about the liquid water?
    d. What is meant by the symbols in brackets?
    e. Is the equation balanced? If not, balance it.
    f. Why is it vital to have a balanced equation?

g. What does the equation say about the ratios in which these reactants combine to form the product?
h. How many moles of hydrogen gas must react with 1 mole of oxygen gas?
i. How many moles of product would one expect?
j. Would it be possible to work out what mass of hydrogen and oxygen is needed for this reaction? How?
k. Would it be possible to work out the mass of product that could be expected? How?
l. Would it be possible to work out what volume of hydrogen and oxygen would be needed for the reaction? How?
m. If you had 4 moles of hydrogen and 1 mole of oxygen how many moles of water would be produced? Explain your reasoning. Which reactant is limiting the amount of product that can be formed?
n. If you had 2 moles of hydrogen and 6 moles of oxygen how much water would be produced? Explain your reasoning. Which reactant is limiting the amount of product that can be formed?

Got it: Stoichiometry is an accounting system for chemical reactions. It tells us how much of each reactant is required to combine completely to form particular products. The information is given in **balanced equations** which relate moles (and therefore numbers of particles) of reactants and products. From the moles relationships one can continue to relate masses, volumes and concentrations of the various reactants and products.

Activity 2.2.12: Test your understanding of the concept of stoichiometry

Consider the equation below (describing a neutralisation reaction between sulfuric acid and sodium hydroxide) and then answer the questions that follow:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

1. How many moles of sulfuric acid do I need to make 5 moles of sodium sulfate?
2. If I have 3 moles of sulfuric acid how many moles of sodium hydroxide are needed to neutralise the acid completely?
3. If I have 3 moles of sulfuric acid what mass of sodium hydroxide is needed to neutralise it completely?
4. If I have 1 mole of sulfuric acid and 20 moles of sodium hydroxide how many moles of sodium sulfate can I make?
5. If I have 1 mole of sulfuric acid and 5 moles of sodium hydroxide how many moles of sodium sulfate can I make?
6. In question 5 would any reactant be left over? Which one? How many moles? What mass?
7. If I mix 1 mole of sulfuric acid and 2 moles of sodium hydroxide what mass of sodium sulfate can I expect?
8. If I do the procedure outlined in 7 and obtain 131.5 g of sodium sulfate what would the percentage yield be?

Answers 2.2.12:

1. 1 mole of H_2SO_4 can give 1 mole of Na_2SO_4
   5 mole of H_2SO_4 can give 5 mole of Na_2SO_4
2. 1 mole of H_2SO_4 requires 2 moles of NaOH
   3 mole of H_2SO_4 requires 6 moles of NaOH
3. 6 moles of NaOH are needed (question 2)
   RMM of NaOH = 22.99 + 16.00 + 1.01 = 40.00 $g.mol^{-1}$
   Mass of 6 moles = 6 x 40.00 g = 240.0 g
4. 1 mole of H_2SO_4 requires 2 moles of NaOH
   20 moles of NaOH is excess NaOH – only 2 moles are required for the acid present. 18 moles of NaOH will be left over.
5. 1 mole of H_2SO_4 requires 2 moles of NaOH.
   NaOH is present in excess – only 2 moles of NaOH are needed.
   1 mole of H_2SO_4 can give 1 mole of Na_2SO_4
6. Yes 5-2 = 3 moles of NaOH will be left over.
   RMM of NaOH = 22.99 + 16.00 + 1.01 = 40.00 $g.mol^{-1}$
   Mass of 3 moles = 3 x 40.00 g = 120.0 g
7. 1 mole of H_2SO_4 and 2 moles of NaOH give 1 mole of Na_2SO_4
   RMM of Na_2SO_4 = (2 x 22.99) + 32.07 + (4 x 16.00) = 142.1 $g.mol^{-1}$
   Mass of 1 mole of Na_2SO_4 = 1 x 142.1 = 142.1 g

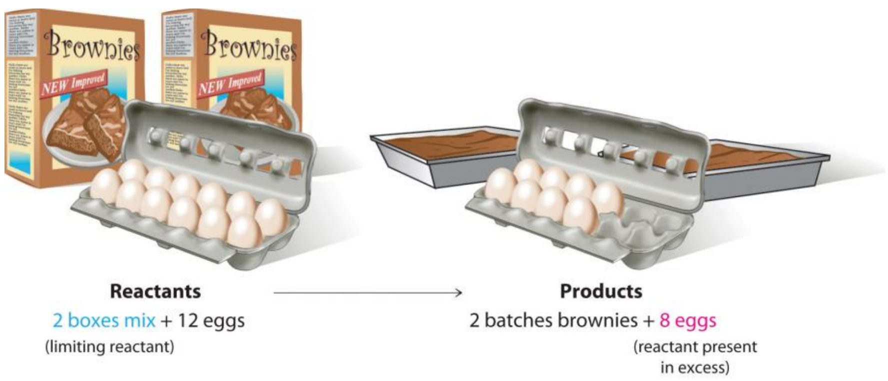
8. % yield = actual mass/theoretical mass x 100%
   = 131.5 g/142.1 g x 100% = 92.5%

Unit 2.2.8 Limiting reagent

While you were working with the reaction above you noticed that sometimes a reagent would be limiting and sometimes it would be present in excess. Look back at Questions 4, 5 and 6 above.

A **limiting reagent** is one that is present in too small an amount to convert the other reagents completely to product.

A **reagent in excess** is one that is present in too high a quantity to be completely converted into product. So some of this reagent would be left over after the reaction stopped taking place.



Unit 2.2.9 Theoretical yield and Percentage yield

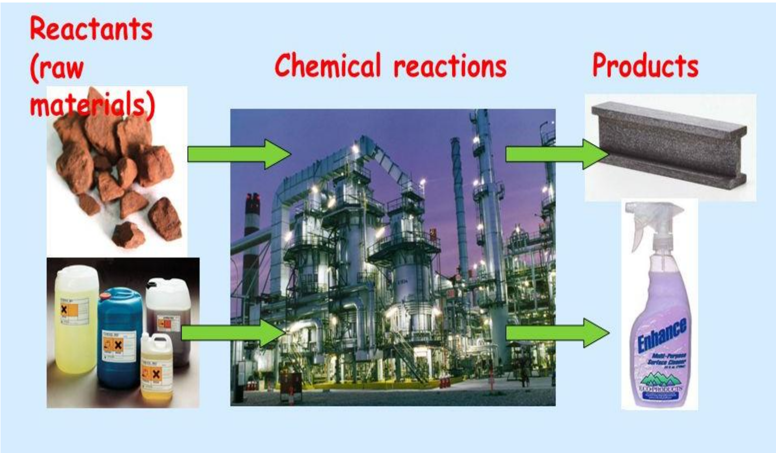
When you have an equation representing a reaction, and you know the amounts of all the reactants present, you can work out how much of a particular product is theoretically possible to make. This is called the theoretical yield. You did this kind of calculation in question 7 above.

In a laboratory, however, reactions seldom work perfectly so the theoretical yield is seldom obtained. When you measure the actual amount of product you are measuring actual yield. This is usually less than the theoretical yield.

To get an idea of how efficient your reaction is you would calculate percentage yield as follows:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

You did such a calculation in question 8 above.



In industry it is important to know what the % yield is because it will determine your profit. If it is low you would find ways to improve the efficiency of your manufacturing process. Chemists who can raise the 5 yield for their employers are highly valued and should be well paid!

Activity 2.2.13: Answer the following questions:

1. Refer to the Periodic Table and answer the following questions:
   a. State the average mass of 1 atom of phosphorus
   b. State the average mass of 1 mole of carbon
   c. State the average mass of 6.02×10^{23} atoms of zinc
   d. State the average mass of 1 atom of uranium
   e. State the average mass of 1 mole of potassium

2. Calculate the number of particles in the following:
   a. 2.3 moles of sodium atoms
   b. 4.13 moles of carbon dioxide molecules

3. Calculate the relative molar mass of the following compounds:
   a. Silver sulfate, Ag_2SO_4
   b. Sulfur hexafluoride, SF_6
   c. Manganese nitrate, $Mn(NO_3)_2$
   d. Dinitrogen pentoxide, N_2O_5

   e. Strontium phosphate, $Sr_3(PO_4)_2$

4. What is the percentage composition of all the elements in the following examples:
   a. iron oxide, Fe_2O_3
   b. octane (in petrol), C_8H_{18}

5. How much iron would you expect to extract from 1000.0 kg of iron oxide, (Fe_2O_3)?

6. What is meant by standard temperature and pressure?

7. Define molar volume.

8. How many molecules of carbon dioxide would you find in 100.0 g of the gas?

9. How many moles of ammonia gas would be found in a container with a volume of 1.00 dm^3 when the temperature is kept at 0^oC and the pressure at one atmosphere?

10. Fill in the blanks: 2.0 moles of K_2SO_4
    a. contains _____ moles of K atoms
    b. contains _____ g of sulfur
    c. contains _____ individual O atoms
    d. contains _____ formula units of the compound
    e. has a mass of _____ g

11. Calculate the molar concentration for each of the following solutions:
    a. 1.75 g of NaCl in 363.5 ml of water
    b. 20.0 g of sugar ($C_{12}H_{22}O_6$) in a cup of tea (250 ml)

12. What mass of sugar would be present in one litre of solution that has a concentration of 2.5 $mol.dm^3$?

13. Consider the equation below and then answer the questions that follow:

$$CH_4(g) \; + \; 2O_2(g) \; \rightarrow \; CO_2(g) \; + \; 2H_2O(g)$$

a. If 3 moles of CH_4(g) was available, how many moles of oxygen
would be needed to combust it completely to carbon dioxide and water? (1)
b. What volume of oxygen gas would be required at STP? (2)
c. If you had 5 moles of CH_4(g) and 7 moles of O_2(g) present, how many moles of CO_2(g) could you expect?
d. If you had 1000.0 g of oxygen available, what mass of CH_4(g) would you be able to convert to CO_2(g)

14. Consider the following equation:

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

If the reaction of 20.7 grams of $CaCO_3$ produces 6.81 grams of CaO, what is the percent yield?

Subtopic 2.3 Energy changes during chemical reactions

At the end of this subtopic you should be able to:
2.3.1 Define exothermic and endothermic;
2.3.2 Label and interpret potential energy profiles for exothermic and endothermic reactions;
2.3.3 Draw labelled potential energy profiles for exothermic and endothermic chemicalreactions;
2.3.4 Explain the effect of catalysis using potential energy profiles.

Any chemical reaction involves the breaking of existing intramolecular bonds and the formation of new bonds. Another way of saying this is that, during a chemical reaction, bonds in the reactants are broken so as to allow the formation of new bonds in the products.

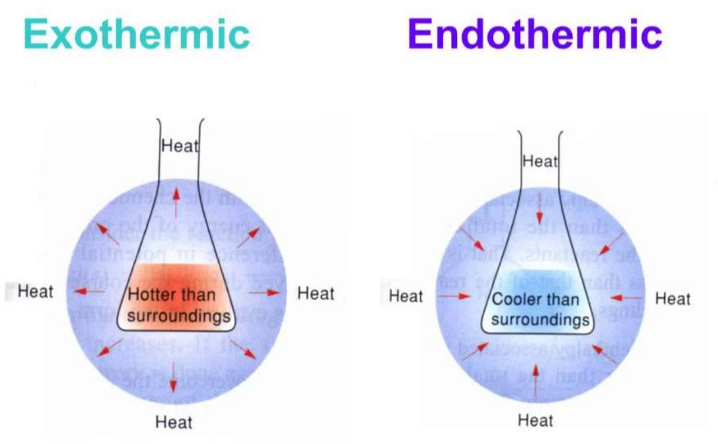
Both the breaking and making of chemical bonds involve the exchange of energy:

In order to break bonds energy is required. Energy-requiring activities are called **endothermic**, so bond breaking is an endothermic process.

When new bonds are formed energy is released. Energy releasing activities are called **exothermic**, so bond making is a exothermic process.

It is possible to do a kind of energy accounting where the energy required to break the bonds of the reactants in a reaction is compared to the energy released during the formation of new bonds in the products. This will let you know whether the whole chemical reaction is exothermic (energy releasing) or endothermic (energy consuming).

So a reaction would be an **exothermic reaction** if more energy is released when the new bonds are being formed than was consumed in the breaking of the bonds.



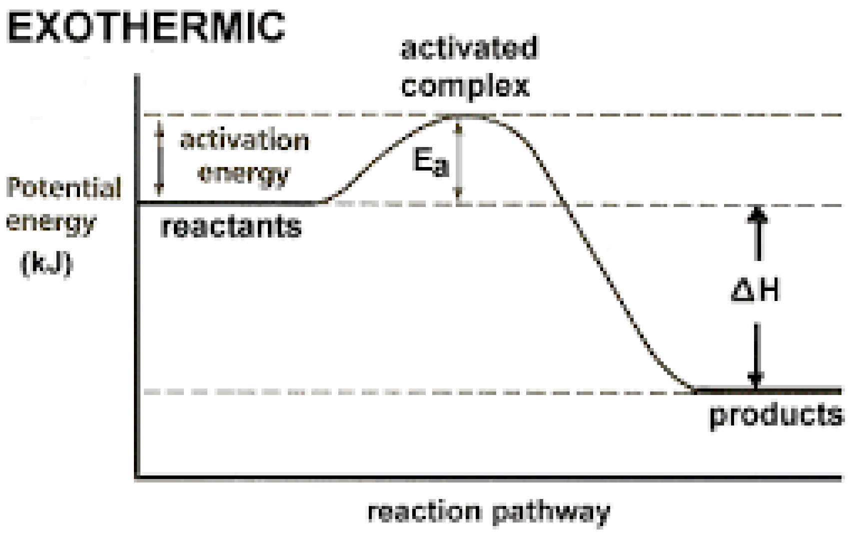
An **endothermic reaction** would be one which requires more energy for the formation of the new bonds than is released in the breaking of the bonds of the reactants.

- If more heat energy is released when making the bonds than was taken in, the reaction is **exothermic**
- If more heat energy was taken in when making the bonds than was released, the reaction is **endothermic**

Unit 2.3.1 Energy diagrams

Energy diagrams (also called potential energy profiles) provide a visual representation of what is taking place during a particular reaction. They show the potential energy levels of the reactants and of the products. The greater the difference between the two, the more energy either released or used up in the reaction.

Consider the following energy diagram for an exothermic reaction:



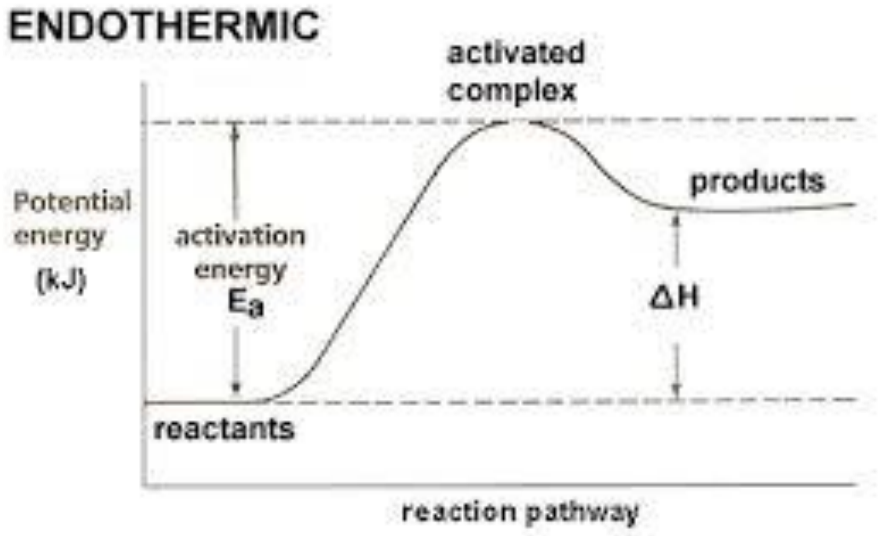
In exothermic reactions the potential energy levels of the reactants are higher than the potential energy present in the products.

The difference between the energy of the reactants and the energy of the products ($H_{products}$ $–H_{reactants}$) is called the **enthalpy change** (ΔH) of the reaction. For an exothermic reaction, the enthalpy change is always negative because the amount of energy present in the products is less than that present in the reactants.

A general equation can be written as follows:

$$\text{Exothermic reactions: } (H_{products} - H_{reactants}) = -\Delta H$$

Now look at an energy diagram for an endothermic reaction:



In an endothermic reaction, the products are at a higher energy than the reactants. This means that the **enthalpy change** of the reaction (ΔH) is positive.

A general equation can be written as follows:

$$\text{Endothermic reactions: } (H_{products} - H_{reactants}) = + \Delta H$$

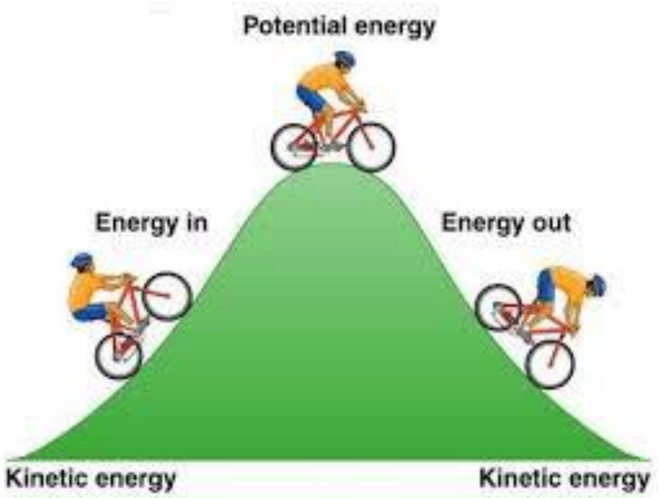
Activation energy (E_a)

You will notice that in both energy diagrams there is a hump in the curve labeled E_a. This is the activation energy.

Activation energy is the minimum energy which must be available reactants for them to be able to be converted into products.

You can get an idea of what activation energy is from the following diagram:

In order for the cyclist to get up the hill he needs to put in a lot of work (energy). However, once he reaches the top of the hill he can free-wheel down the other side.

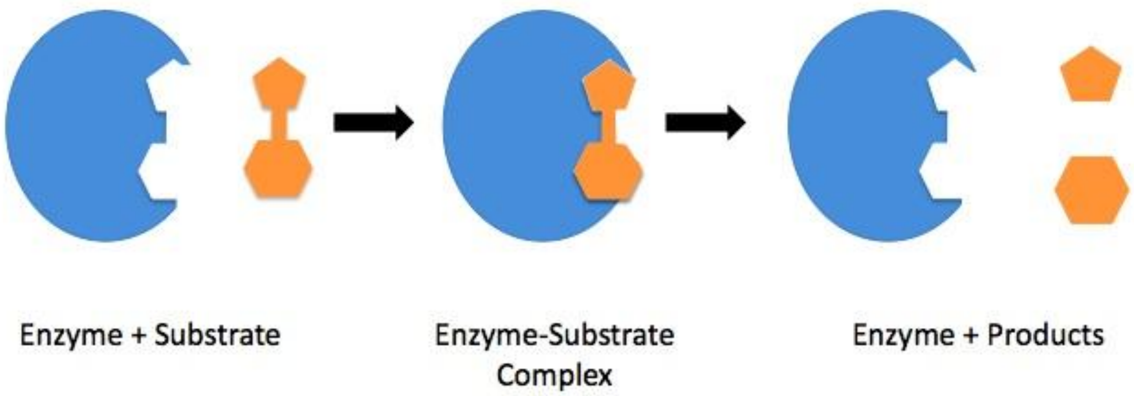


https://www.youtube.com/watch?v=YacsIU97OFc

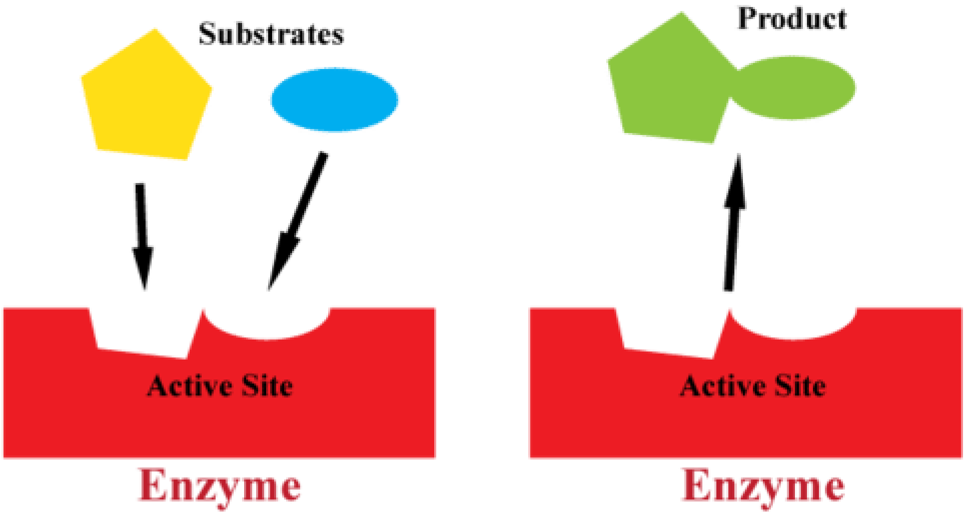
Unit 2.3.2 Catalysts

The rate of a reaction can be increased by adding a suitable **catalyst**. A catalyst is a substance which changes the rate of reaction but is unchanged at the end of the reaction.

The diagram gives you an idea of how a particular type of catalyst, called an enzyme, works: it binds to the reactant in such a way that the bonds are more easily broken. The products are then released, leaving the enzyme ready to repeat the process on another reactant molecule.



Another example, where two reactants combine to form a new product:

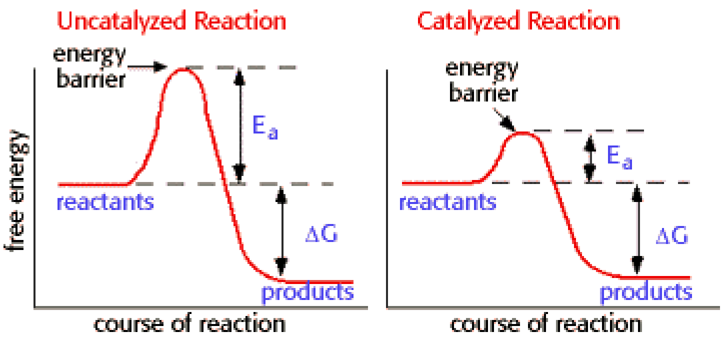


Only a very small amount of catalyst is needed to increase the rate of reaction between large amounts of reactants. This is because the enzyme is constantly released to repeat the task once the product has formed.

A catalyst is specific to a particular reaction:

- Different catalysts catalyse different reactions
- Not all reactions have suitable catalysts
- Biological reactions are catalysed by special types of catalysts called an enzymes

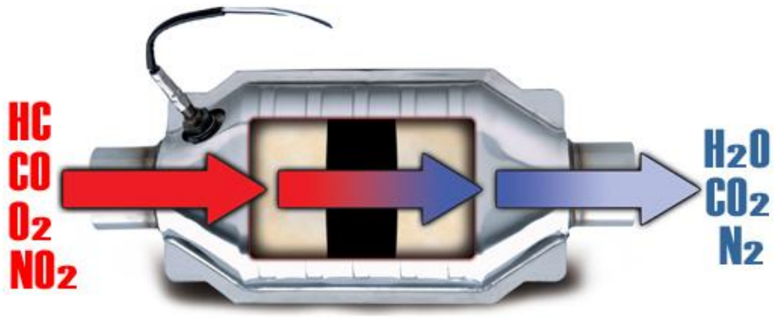
The effect of a catalyst can be shown on and energy diagram as follows:



You will notice that the free energy of the reactants and the products remains the same in both reactions. The only thing that changes is E_a, the size of the energy barrier that has to be overcome for the reaction to take place. It is lowered in the presence of the enzyme.

Did you know:

To reduce the production of toxic gases in the exhaust fumes of cars, modern cars have a catalytic converter. Catalytic converters speed up the reactions between carbon monoxide and



unburned fuel with oxygen from the air to form less harmful carbon dioxide and water. They use catalysts made of platinum and rhodium. These catalysts work best at the higher temperatures found in an engine.

Some types of catalysts:

- Sulfuric acid acts as a catalyst in substitution reactions such as when an organic alcohol is converted into an organic alkyl halide.

$$CH_3CH_2OH(g) + HCl(g) \xrightarrow{H_2SO_4} CH_3CH_2Cl + H_2O(\ell)$$

- Hydrogen peroxide catalyses the decomposition reaction of water into hydrogen and oxygen gases.

$$2H_2O(l) \xrightarrow{H_2O_2} 2H_2(g) + O_2(g)$$

- Enzymes
  The names of enzymes usually end in the letters **-ase**. Three of the most common families of enzymes involved in digestion are:
  - **lipases** – these break down fats
  - **proteases** - break down proteins
  - **carbohydrases** - break down carbohydrates.

Activity 2.2.14:

1. A catalyst increases the rate of a reaction by

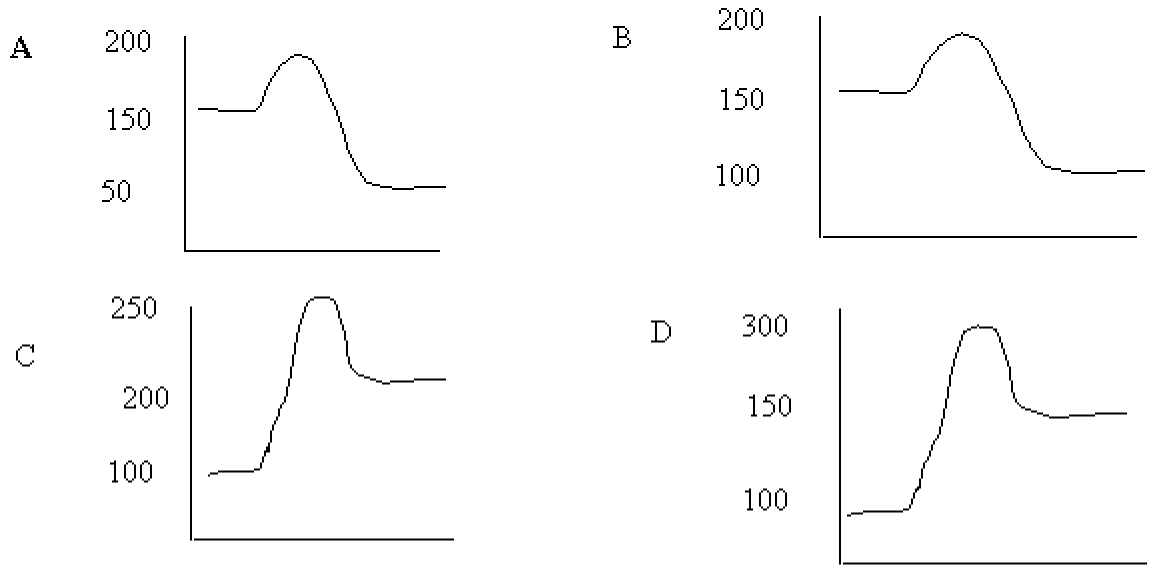
   A. Increasing the concentration of the reactant(s)

   B. Decreasing the concentration of the reactant(s)

C. Increasing the activation energy of the overall reaction

D. Decreasing the activation energy of the overall reaction

2. A forward reaction has an activation energy of 50 kJ and a ΔH of –100 kJ.



Which energy diagram most accurately describes this reaction?

3. What is the activation energy of a reaction, and how is this energy related to the activated complex of the reaction?

4. What happens when a catalyst is used in a reaction?

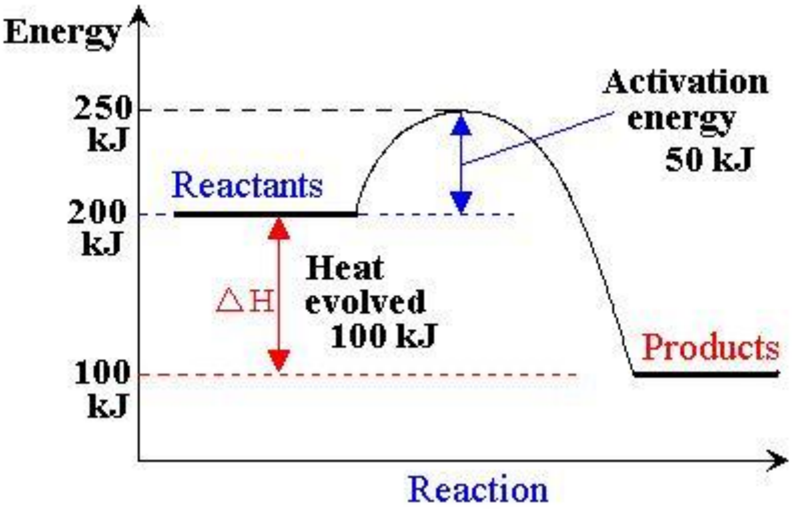
5. Name 4 things that will speed up or slow down a chemical reaction.

6. Draw an energy diagram for a reaction.
   Label the axes.
   Potential energy(PE) of reactants = 350 KJ/mol,
   Ea = 100 KJ/mol,
   PE of products = 250 KJ/mol.

   Is this reaction in exothermic or endothermic? Explain.

7. How could you lower the activation energy for the reaction in #6?

1. The reaction is exothermic. The △H is -100 KJ/mol which means heat is released.



2. Add a catalyst

Subtopic 2.5 Chemical equilibrium

At the end of this subtopic you should be able to:
2.5.1 Describe what is meant by a reversible reaction;
2.5.2 Draw simple graphs of concentration against time and reaction rate against time to illustrate chemical equilibrium (only include graphs that illustrate the process of the initial equilibrium, exclude the impact of various factors that alter the position of the equilibrium);
2.5.3 State Le Chatelier's principle;
2.5.4 List the factors that affect equilibrium, namely temperature, concentration and pressure (for gases);
2.5.5 Apply Le Chatelier's principle to explain the effect of changes in the factors that affect equilibrium.

Initially chemists believed that all reactions were irreversible. This means that reactants turn into products only – products are not able to return to the form of the reactants.

Unit 2.5.1 Irreversible and reversible reactions

Irreversible reactions

Chemical equations for irreversible reactions are written as follows:

$$A + B \rightarrow C + D$$

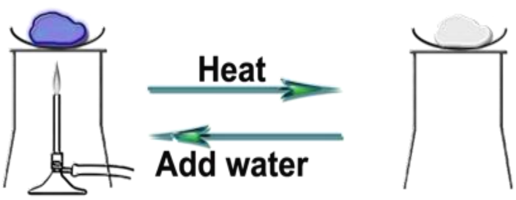
A and B represent reactants and C and D represent products. The reaction direction (shown by the arrow) goes only in one direction from reactants to products.

Many reactions are of this type. For example:

- When you bake a cake, several ingredients are combined together and heated to produce something new and delicious. But you cannot get back to the original ingredients again.
- Burning fuel is not reversible. These kinds of reactions are called combustion reactions. An organic compound reacts with oxygen to produce carbon dioxide and water.

Reversible reactions

In 1803, a French chemist called Claude Louis Berthollet realized that some reactions are in fact able to return from products to reactants. These are called reversible reactions. Most reactions are not reversible (they are irreversible). However, there are several reversible reactions that are very important industrially.



Chemical equations for reversible reactions are written as follows:

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

Again, A and B represent reactants and C and D represent products. The reaction direction (shown by the arrow) goes in both directions, from reactants to products or from products to reactants.

Got it: A reversible reaction is shown by the sign, $\rightleftharpoons$

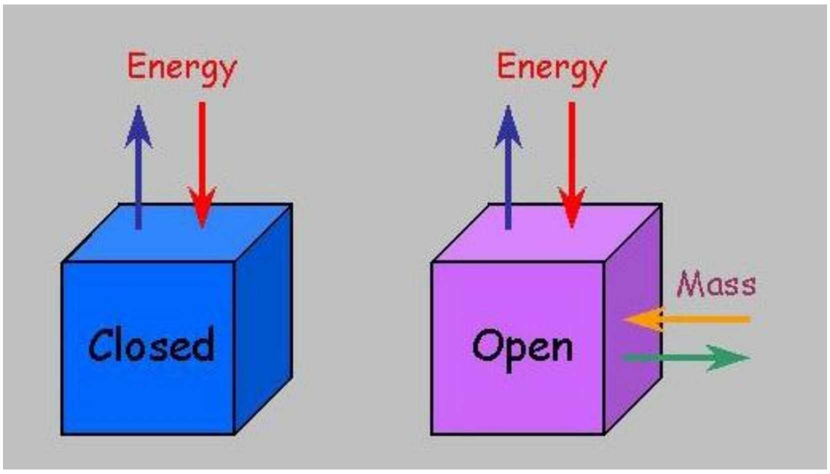
- a half-arrow to the right shows the direction of the **forward reaction.** The forward reaction is always the reaction going from left to right.
- a half-arrow to the left shows the direction of the **backward reaction.** The backward reaction is always the reaction going from right to left.

Unit 2.5.2 Dynamic equilibrium

When a reversible reaction is allowed to continue uninterrupted, there will come a time when the rate at which the forward reaction is taking place equals the rate at which the reverse reaction is taking place. When this happens no further overall change can be measured. In fact, it looks as though the reaction has stopped altogether. This is not actually the case. No overall change is measurable, but the forward and back reactions are continuing all the time, at the same rate.

The important thing to remember is that both the forward and the reverse reactions are continually happening, but the **concentrations (amounts) of each reactant and each product remain constant.**

Closed systems



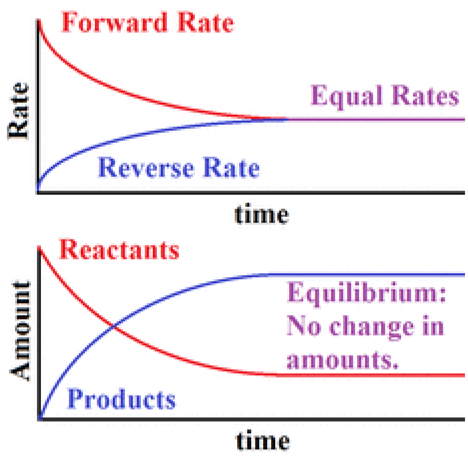
It is only possible to measure a reaction in dynamic equilibrium in a closed system. In a closed system the reactants are placed into a container that can be shut off from the air. In such a system no reactants or products can escape from the container or be added to the container.

If the reaction is a reversible one, as you now know, two possible reactions can take place here, one where A + B is converted to C + D and a second where C + D is converted to A + B. What happens if these are taking place at the same time in a closed system?

Actually, neither the reactants nor the products are ever used up completely because they are both constantly and simultaneously (at the same time) being produced.

If a reversible reaction takes place in a closed system it will eventually reach a state of dynamic **equilibrium**.

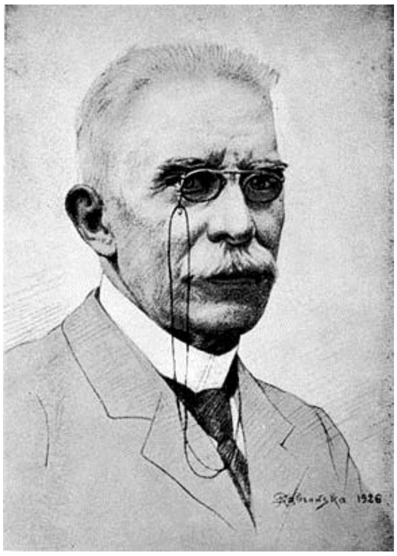
At equilibrium the rates of the forward and reverse reactions are the same. In addition, the amounts of reactants and products in the container remain the same as well.



Got it: A **reversible** reaction is a chemical reaction in which the products can be converted back to the original reactants under suitable conditions.

When a reversible reaction reaches equilibrium, the forward and reverse reaction rates are the same and the amounts of reactants and products remain the same.

Unit 2.5.3 le Chatelier’s Principle



This principle was established in 1884 by the French chemist and engineer Henry-Louis Le Chatelier. The principle states:

If a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium shifts to counteract the change to reestablish an equilibrium

Le Chatelier's principle describes what happens when something disturbs a system that is in equlilibrium. Here you will consider three ways in which you can change the conditions of a chemical reaction at equilibrium:

- changing the concentration of one of the components of the reaction
- changing the pressure on the system
- changing the temperature at which the reaction is run.

By making changes such as these it is possible to force a reversible reaction to favour one direction or the other.

Unit 2.5.4 Concentration changes

Consider the following reversible reaction:

$$A + 2B \rightleftharpoons C + D$$

What do you think happens the concentration of A is increased?

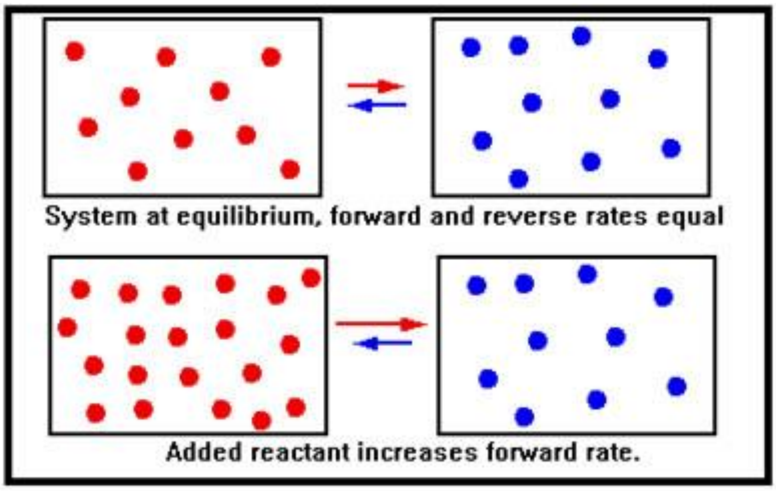
According to Le Châtelier, the position of equilibrium will move to counteract the change. In order to decrease the raised concentration of A, the A will react with B to produce more C and D. The equilibrium moves to the right.

This kind of intervention would be valuable if B was relatively expensive reactant in the formation of the products C and D. So, in order to maximize the amount of B converted into product one could raise the concentration of A.

What do you think happens the concentration of A is decreased?

In order to make more of A, the reverse reaction would be favoured using up more of C and D in the process.

Usually the concentration of a product is decreased. So, if C is removed from the system the reaction will move to the right to re-establish the equilibrium. In this way more of A and B are converted into product.



Summary of concentration effects:

- Add more reactant – shift towards products
- Remove reactants – shift towards reactants
- Add more products – shift towards reactants
- Remove products – shift towards products

Unit 2.5.5 Pressure changes

Note: Pressure changes only apply to reversible reactions involving gases (or at least one gas).

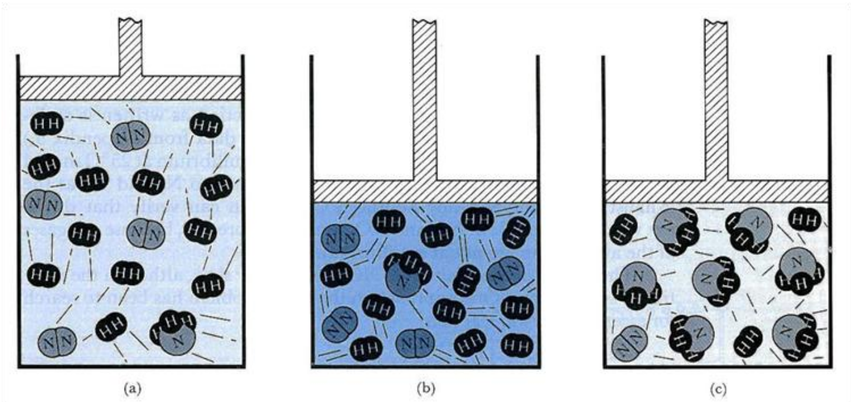
Increasing the pressure

Consider the following reversible reaction:

$$A(g) + 2B(g) \rightleftharpoons C(g) + D(g)$$

According to Le Châtelier, if the pressure is increased, the position of equilibrium will move to reduce the pressure again. Pressure is caused by gas molecules hitting the sides of their container. The more molecules (moles of gas) in the container, the higher the pressure will be. The system can reduce the pressure by reacting in such a way as to produce fewer molecules.

In this case, there are three moles of reactants, but only two moles of products. So, by forming more C and D, the system causes the pressure to reduce. Increasing the pressure on a gas reaction shifts the position of equilibrium towards the side with fewer moles of gas molecules.



(a) 17 gaseous reactant molecules at equilibrium
(b) Same 17 gaseous molecules now under increased pressure.
(c) Shift towards more products takes place as this reduces the number of gaseous particles to 11. A new equilibrium is established.

Decreasing the pressure

The equilibrium will move to increase the pressure again. This happens when more gas molecules are produced. In this example that means the reaction would have to move to the left.

What do you think would happen if the pressure were changed in a reaction like the following?

$$2A(g) + 3B(g) \rightleftharpoons 4C(g) + D(g)$$

Since the number of molecules (moles) is the same on both sides of the equation (5) a pressure change will have no effect on the position of the equilibrium.

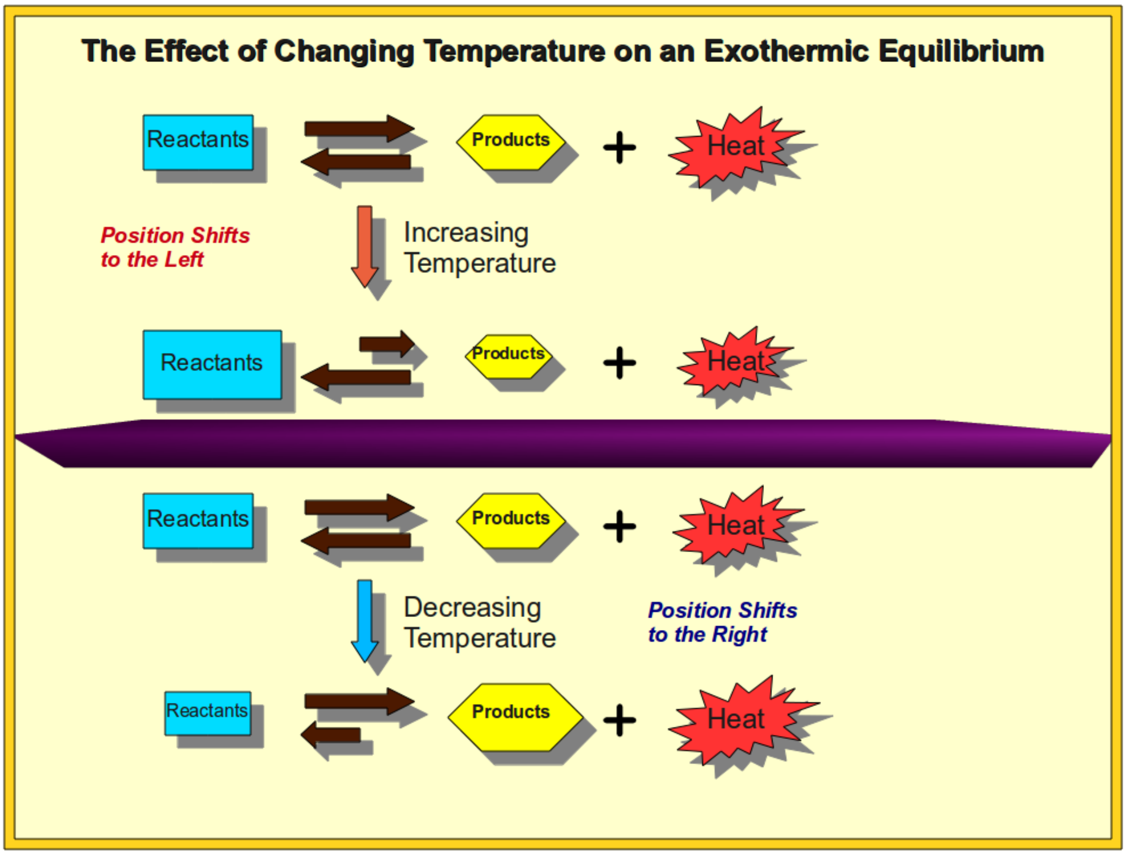
Summary of Pressure Effects

- By adding or removing a gaseous reactant or product the equilibrium will move in the direction of the smaller number of molecules (moles) present.
- By increasing the pressure in the system (usually by decreasing the volume of the container) the equilibrium will move towards the side with the lower number of molecules (moles)
- By decreasing the pressure in the system (usually by increasing the volume of the container) the equilibrium will move towards the side with the higher number of molecules (moles)

2.5.6 Temperature changes

Increasing the temperature

If the temperature is increased, then the position of equilibrium will move so that the temperature is reduced again. You already know about exothermic and endothermic reactions. If the forward reaction is exothermic (releasing heat energy), then the reverse reaction will be endothermic (using up heat energy).



Consider the following:

$$A + 2B \rightleftharpoons C + D \quad \Delta H = -250\ kJ\ mol^{-1}$$

(A negative ΔH value tells you that the forward reaction is exothermic and heat is give out).

This system is in equilibrium at 350°C, and the temperature is increased 525°C. The forward reaction is exothermic. To reduce the overall temperature, it needs to absorb the extra heat that has been added.

Therefore, in order to absorb heat from the system the reaction will have to favour the endothermic reaction which is the reverse reaction. So at the new equilibrium point there will be more reactants (A and B) present.

Ask yourself: Would you increase the temperature of the reaction if you had a factory producing C and D for sale in the example just discussed?

Decreasing the temperature

If the temperature is decreased, then the position of equilibrium will move so that the temperature is increased again. So, in the example above a decrease in temperature will push the reaction to the left hand side (the reverse reaction) and a new equilibrium will be established.

Summary of the effects of temperature

- Increasing the temperature of a system in dynamic equilibrium favours the endothermic reaction. The system reacts to the change by absorbing the extra heat which it achieves by reversing the direction of the reaction.
- Decreasing the temperature of a system in dynamic equilibrium favours the exothermic reaction. The system reacts to the change by releasing more heat which it achieves by reversing the direction of the reaction.

Unit 2.5.7 Catalysts

Remember that catalysts do not take part in chemical reactions – they simply change the rate at which both the forward and the reverse reaction takes place.

So, adding a catalyst to a reaction in dynamic equilibrium will make no difference to the position of the equilibrium.

What a catalyst can do, however, is to speed up the time it takes for the equilibrium position to be reached.

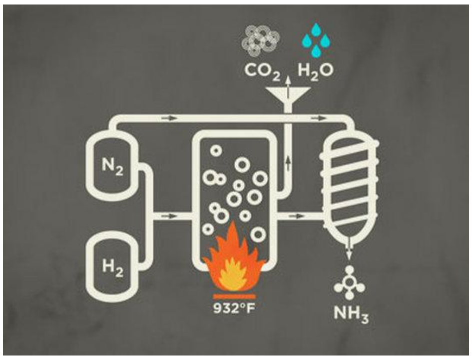
Unit 2.5.8 Forcing a reversible reaction in one direction or the other

Sometimes, especially in industry, it is required to make sure that as much of the reactants is converted into products as possible. This is because the product can be sold and money earned.

As you have already learned, this can be achieved by changing the reaction conditions in a number of ways:

- The concentration of the reactants (or products) can be changed. For example, more reactants can be put into the system and products can be removed from the system.

- The pressure in the reaction container can be increased or decreased.
- The temperature at which the reaction takes place can be increased or decreased



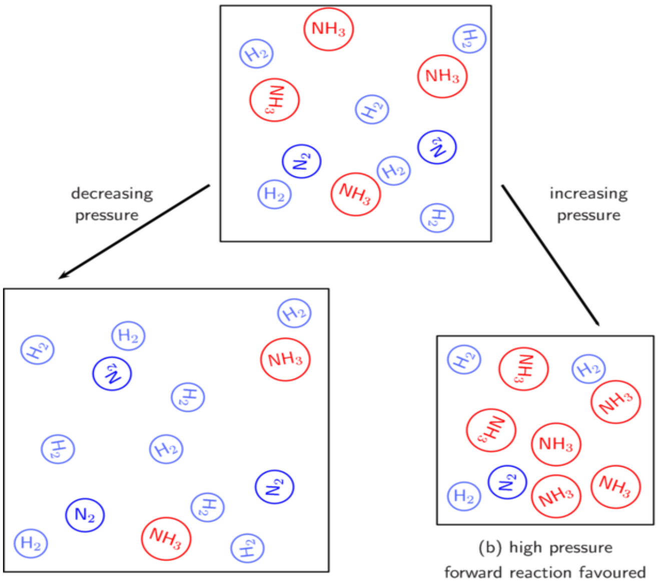
A useful example of a reversible reaction is as follows:

The synthesis of ammonia from nitrogen and hydrogen (The Haber Process in industry)

$N_2(g) + 3H_2(g) \rightleftharpoons 3NH_3(g)$

By changing the temperature and pressure of the reacting gases you can make the reaction go one way more than another

- High pressure forces the reaction to go the right (because there are fewer moles of $NH_3(g)$ (3) than moles of $N_2(g)$ + $H_2(g)$ (4).
- Higher temperature forces the reaction to the right as the ammonia decomposes into nitrogen and hydrogen gases again.
- Therefore, in order to maximize the production of ammonia gas the pressure needs to be increased and the temperature lowered.



(a) low pressure
reverse reaction favoured

(b) high pressure
forward reaction favoured

Some other examples of reversible reactions:

1. Thermal (using heat) decomposition of ammonium chloride

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

2. Dehydration of blue copper sulfate crystals to form white anhydrous copper sulfate

$$CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(l)$$

3. Reaction between sulfur dioxide and oxygen to form sulfur trioxide

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Activity 2.2.15: Understanding Le Chatelier's Principle and Dynamic Equilibria

1. Consider the given equation of a chemical reaction in dynamic equilibrium and answer the questions that follow, remembering what Le Chatelier's Principle says about changes in such reactions:

$$2A + 3B \rightleftharpoons C + 2D \quad -\Delta H$$

   a. What would happen to the equilibrium if one increased the concentration of A in the reaction?
      Hint: What could the reaction do to decrease the concentration of A again?
   b. What would happen to the equilibrium if one decreased the concentration of A in the reaction?
      Hint: What could the reaction do to increase the concentration of A again?
   c. If the reaction given above involved gases, what would happen to the equilibrium if the pressure was increased?
   d. If the reaction given above involved gases, what would happen to the equilibrium if the pressure was decreased?
   e. What would happen to the equilibrium in a reaction where both sides of the equation have the same number of gaseous particles (e.g. $A + B \rightleftharpoons C + D$) and the pressure was changed?
   f. What would happen to the equilibrium of the reaction if the temperature was increased?

Remember that -ΔH indicates that the forward reaction is exothermic i.e. gives out heat, and the reverse reaction is endothermic i.e. takes in heat. The amount of heat given out is exactly equal to the amount of heat taken in.

g. What would happen to the equilibrium of the reaction if the temperature was decreased?
h. What effect might a catalyst have of the equilibrium of a reaction?
i. If you wanted to make more C and D in the present example what changes would you make to the reaction conditions?

2. Consider the reaction for the synthesis of ammonia from nitrogen and hydrogen, then answer the questions that follow:

$$N_2(g) \quad + \quad H_2(g) \quad \rightleftharpoons \quad NH_3(g)$$

a. Write a balanced equation for the reaction
b. If you were interested in the commercial manufacture of ammonia in which direction would you like the equilibrium to be shifted?
c. Will more ammonia form if the temperature is raised? Explain your answer. Note: The forward reaction is exothermic
d. Can you suggest what might happen if you increased the pressure of the reaction? Hint: The greater the number of particles present in gases, the greater the pressure.
e. What would happen if you removed ammonia from the reaction vessel as it was being formed?
f. What would happen if a catalyst was added to the system? The commonly used catalyst for this reaction is iron.
g. Imagine that you have a company that produces ammonia for sale. What conditions would you suggest to maximise your ammonia output and therefore your profit?

Answers 2.2.15

Activity 2.2.16

1. Explain what is meant by a reaction that is in dynamic equilibrium.

2. State Le Chatelier's Principle

3. For a reversible reaction between gases, if the forward reaction is exothermic what is the effect of raising the temperature?

4. For a reversible reaction between gases, if the forward reaction is endothermic what is the effect of raising the temperature?

5. Consider the following reaction which describes what happens in a catalytic converter used to reduce carbon monoxide emissions in motor car exhausts:
   $$2CO(g) + O_2(g) \leftrightarrow 2CO_2(g)$$
   Suppose the reaction is taking place in a closed vessel and is at equilibrium.
   a. What would happen to the reaction equilibrium and the concentration of CO if a platinum catalyst were added?

   b. What would happen to the reaction equilibrium and the concentration of CO if the temperature were increased?

      (Note: this is an exothermic reaction in the forward direction).
   c. What would happen to the reaction equilibrium and the concentration of CO if the pressure were increased?

   d. What would happen to the reaction equilibrium and the concentration of CO if the concentration of oxygen were increased?
6. Name the industrial process for the production of ammonia and write down the complete, balanced equation for its formation.

Answers 2.2.16:

1. A reaction that is in dynamic equilibrium is one in which the rate of the forward reaction and the rate of the reverse reaction are equal, so the concentrations of reactants and products no longer change.

2. Le Chatelier's Principle states that if a reversible reaction that is in a state of dynamic equilibrium is disturbed by changing the conditions, the reaction will do what it can to oppose that change.

3. For a reversible reaction between gases, if the forward reaction is exothermic raising the temperature?

4. For a reversible reaction between gases, if the forward reaction is endothermic raising the temperature increases the yield.

5.
    a. Equilibrium remains the same and CO concentration does not change

    b. Equilibrium moves to the left and CO concentration increases
    c. Equilibrium moves to the right and CO concentration decreases
    d. Equilibrium moves to the right and CO concentration decreases
6. The Haber Process

    $N_2(g) \quad + \quad 3H_2(g) \quad \leftrightarrow \quad 2NH_3(g)$

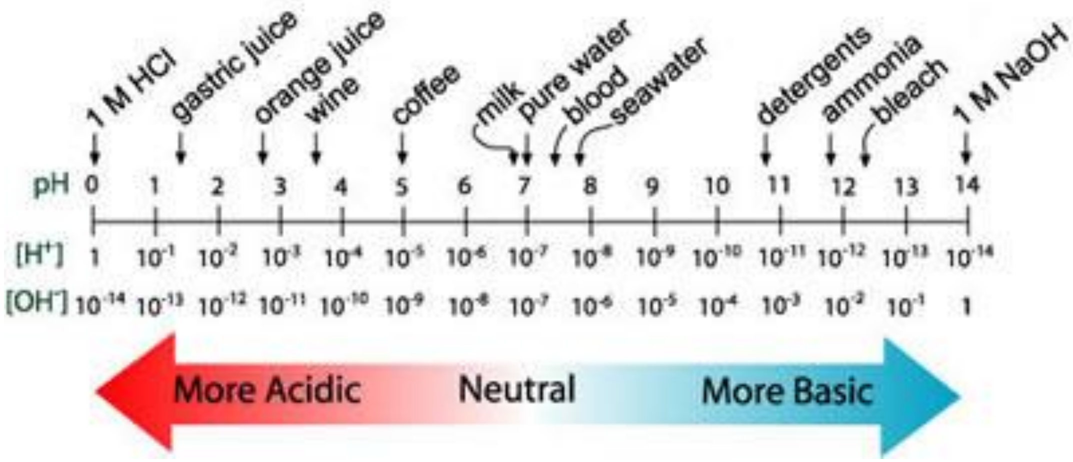
Subtopic 2.6 Acids, bases and neutralization reactions

All substances are either **acidic, basic** or **neutral.** Acids and bases are some of the most common chemicals in both a laboratory and in our everyday lives. For example, vinegar and lemon juice are **acids** that we all know very well. Household ammonia is a **base** that we are familiar with mostly because of its pungent smell. Bases which are soluble in water are called **alkalis**. We can place many chemicals in one of these two groupings according to some rules that allow us to identify an acid and a base. These rules are presented in the following table. Chemicals which are neither acidic nor alkaline are called **neutral**.

Acids	Bases/Alkalis
Acids taste sour e.g. lemon juice and vinegar	Bases have a bitter taste

Some acids are corrosive – they can kill living cells or stop them from working properly	Some bases can damage the eyes and the skin e.g. sodium hydroxide
Some acids are essential to health e.g. vitamin C is ascorbic acid	Bases have a slippery feel (because they change the fats in the skin to soaps!)
Acids react with metals (corrode metals) and give off hydrogen gas and a salt	Some bases are helpful e.g. milk of Magnesia as a laxative or to counteract acidity in the stomach.
Acids react with bases and become neutralised	Bases change the colour of indicators e.g. litmus paper turns blue
Acids react with carbonates and carbon dioxide is given off	Bases have pH values greater than 7
Acids change the colour of indicators (and other dyes) e.g. litmus paper turns pink	Bases react with ammonium compounds to give ammonia gas, water and a salt.
Acids have a pH of less than 7	Bases have a pH above 7
Acids react with metal oxides, metal hydroxides and ammonia solution to give water and a salt	

Fig : The pH scale



Got it:

- The pH scale is a measure of the acidity or alkalinity of a substance.
- A neutral substance has a pH of 7.

- Basic substances have pH values greater than 7 (the higher the number the more strong the alkalinity).
- Acidic substances have pH values below 7 (the stronger the acid the lower the number).
- pH measures the hydrogen ion concentration.
- Indicators are also measures of pH – different colours indicating different hydrogen ion concentrations

Activity 2.2.17: Examining acids and bases

A student tested a number of substances in the laboratory. Each chemical was tested for its solubility (could it dissolve in water?). The resulting solutions were then tested by dipping a strip of litmus paper into the solution. Litmus paper is a rough measure of whether a substance is acidic or alkali. Acidic solutions turn litmus paper red/pink and alkaline solutions turn litmus paper blue.

The results were recorded as follows:

Substance	Behaviour in water (soluble or insoluble)	Effect on red litmus paper	Effect on blue litmus paper	Classification (acidic, basic or neutral)
Sulfuric acid	Soluble	No effect	No effect	Neutral
Table salt	Slightly soluble	No effect	No effect	Neutral
Sodium hydroxide	Soluble	Turns blue	Remains blue	Basic (alkaline)
Hydrochloric acid	Soluble	Remains red	Turns red	Acidic
Potassium hydroxide	Soluble	Turns blue	Remains blue	Basic (alkaline)

1. If you can, write down the chemical formulae of the pure chemicals that you are using. Put the acids and alkalis into separate groups.

2. What common element do you notice for all the acids? What common ion would be produced when an acidic substance is placed in water?
3. What do you notice to be common in the formulae for the alkalis? What common ion would be produced when a basic substance is dissolved in water?
4. The student then performed the following experiment:
   A small volume of the acid, HCl, was placed into a test tube and tested with blue and red litmus papers. Then a drop of the alkaline NaOH solution was added to the test tube , mixed and tested with litmus paper. This process was repeated, adding one drop more NaOH at a time. The student continued doing this until a change in the behaviour of the litmus paper was observed..
    a. Write down a balanced equation for the HCl and NaOH reaction. What do you see? How can you use this information to explain the student's observations?

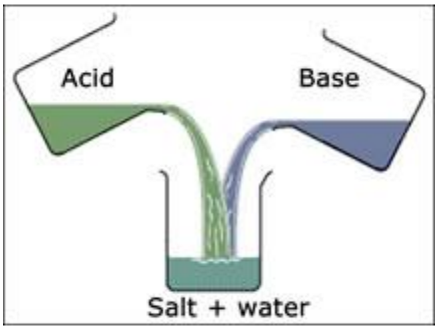
Unit 2.6.1 Reactions between acids and bases – Neutralisation reactions

Knowing what you do about acids and alkalis, what would you expect to see in a neutralisation reaction? Would you expect the products to be acidic, alkaline or neutral? What would it mean to be neutral?

As you have probably already realised, this type of reaction typically results in two neutral products: a salt and water.

A salt is an ionic compound that does not involve $–OH^{-1}$, O^{2-} or H^{+} ions. The salt is composed of the anion from the acid and the cation from the base.

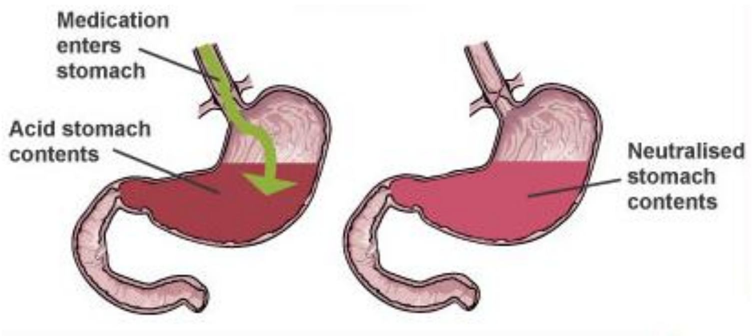
Remember: A cation is a positively charged ion and an anion is a negatively charged ion.



Unit 2.6.2 Neutralisation reaction in everyday life

Neutralisation reactions are often used to deal with situations that occur in our daily lives. Here are some examples:

1. In digestion



Your stomach contains HCl which is essential for digestion. However, if there is too much HCl present it can cause indigestion and even ulcers. So antacids are used to neutralize the stomach acid and relieve the pain and discomfort.

Did you know: Antacids have been around for as long as people have felt the need to deal with stomach discomfort. Originally, herbs and other plants and roots were used, and, in fact, they still are today. For example, ginger root, the liquorice plant, peppermint leaves and chamomile are often used. In South Africa the gel from the leaves of a plant called *Aloe arborescens* (inKalane or Krantz aloe) is still used with good effect.

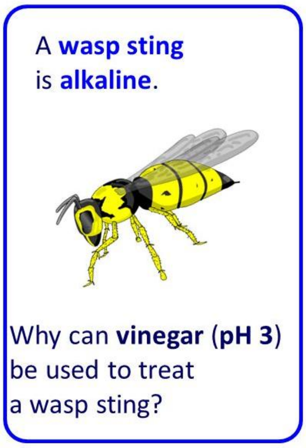
The clinical word for acidity of the stomach is dyspepsia which gives you a clue for the original purpose of peppermints!

Other effective antacids include the sodium bicarbonate and calcium bicarbonate which act as simple acid-base neutralizations in the stomach.

2. Treating bee or ant stings

Insect stings often contain formic acid which is injected into your skin when you are stung or bitten. This causes pain. The pain can be eased by rubbing the area with moist baking soda ($NaHCO_3$, sodium bicarbonate) to neutralize the formic acid. Another treatment is calamine lotion which contains zinc carbonate ($ZnCO_3$) which also neutralises the acid.

Did you know: Many creatures use acids as part of their defence mechanisms and also for collecting food. For example, have you ever been bitten by an ant and felt the sting? That is because the ant injects formic acid into the wound. The amount of formic acid injected is very small, so the pain in a human being is limited, but in another insect such as a termite it is enough to immobilise it or kill it. Why do you think that we can use sodium bicarbonate to treat certain insect stings and bites? Wasps actually inject alkali into the skin when they sting you, so sodium bicarbonate would not be an effective antidote andan acid like vinegar would have to be used!



3. Treating acidic soil in agriculture.
   Excessive use of fertilizers makes soils acidic, resulting in poor growth of crops. If a base such as calcium oxide (CaO, quick lime) or calcium hydroxide ($Ca(OH)_2$, slaked lime) is added the acidity can be neutralised and the soil is again more suitable for the plants.

   Soil can also become too basic in which case acid must be added to correct the pH and make it suitable for crops to grow.

   You can test whether soil is acidic or basic using common household substances.

4. Treating factory waste.

Much factory waste is very acidic. So, if this waste finds its way into our water systems it acidifies the water and makes it unsuitable for human, animals and plants. Such waste needs to be neutralised before being disposed of – but not into the water systems!

http://www.fp.utm.my/projek/psm/webtlr/Neutralisation/learning2a.html
http://www.fp.utm.my/projek/psm/webtlr/Neutralisation/learning2b.html

Activity 2.2 18: Write balanced equations for the following reactions. Decide whether they are neutralization reactions or not. Which products are salts.

1. Nitric acid (HNO_3) and potassium hydroxide (KOH)
2. Zinc (Zn) and copper sulfate ($CuSO_4$)
3. Sulfuric acid (H_2SO_4) and sodium hydroxide (NaOH)
4. Hydrochloric acid (HCl) and magnesium hydroxide ($Mg(OH)_2$)
5. Methane (CH_4) and oxygen (O_2)

Activity 2.2.19 : Answer the following questions:

1. Define the following terms:
    a. An acid
    b. A base
    c. An alkali
    d. A salt
    e. A neutral solution

2. Explain the following:
    a. How to test whether a substance is basic, acidic or neutral.

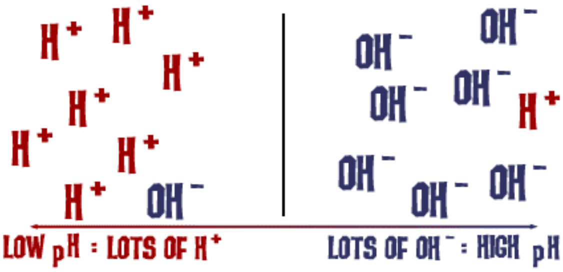
b. What happens when an alkaline solution is slowly added to an acidic solution.

Unit 2.6.3 The Arrhenius and Brønsted-Lowry theories of acids and bases

The Arrhenius theory is based on the following:

- Acids are substances that produce hydrogen ions in solution.
- Bases are substances that produce hydroxide ions in solution.
- Neutralization happens because hydrogen ions and hydroxide ions react to produce water.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$



The examples you have looked at so far described the reactions of Arrhenius acids and bases.

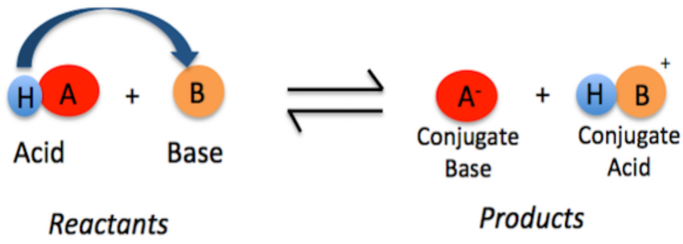
However, there is another theory of acids and bases.

The Brønsted-Lowry theory is based on the following

- Acids are a proton (hydrogen ion, H^+) donors.

- Bases are proton (hydrogen ion, H^+) acceptors.

The Brønsted-Lowry theory does not go against the Arrhenius theory in any way - it just broadens it. Hydroxide ions are still bases because they accept hydrogen ions from acids and form water. An acid produces hydrogen ions in solution because it reacts with the water molecules by giving a proton to them.



Bronsted-Lowry Acid-Base Reaction

Consider the following reaction:

$$HCl(aq) \;+\; H_2O(l) \;\rightarrow\; H_3O^+(aq) \;+\; Cl^-(aq)$$

- According to Arrhenius, hydrochloric acid is an acid because it produces hydrogen ions (H^+) in water
  The hydrogen ions in water become hydroxonium ions (H_3O^+).
- According to the Brønsted-Lowry theory hydrochloric acid is an acid because it is a proton donor.
  A proton is a hydrogen ion (H^+).
  A proton donor is a substance which gives a hydrogen ion away.
  If you look at the reaction above hydrochloric acid gives a hydrogen ion to water.

In the Arrhenius theory the base is a substance that produces -OH^- ions when dissolved in water, as you saw with $NaOH \rightarrow Na^+ + OH^-$.

Remember: You can identify an Arrhenius acid because it contains H atoms which can form H^+ ions. And an Arrhenius base contains OH atoms which can form OH^- ions.

A Brønsted base is a proton acceptor.
This means that a base will gain a hydrogen ion (H^+).

Water acts as a base when it is added to hydrochloric acid because water will gain a hydrogen ion to become H_3O^+.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

$$\text{acid} + \text{base} \rightarrow \text{acid} + \text{base}$$

On the right side of the arrow, H_3O^+ is an acid because it can give away a hydrogen ion to become H_2O.
Cl^- is a base because it can gain a hydrogen ion to become HCl.

Ask yourself: Is water always a base according to Brønsted-Lowry?

Let's look at the Brønsted-Lowry theory a bit closer. Consider the following reaction

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Ammonia (NH_3) is the base because it gains a hydrogen ion (H^+) to become an ammonium ion (NH_4^+).
Water is the acid because it gives away a hydrogen ion (H^+) to ammonia and becomes a hydroxide ion (OH^-).

Now we can see that the reaction can be described like this:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$\text{base} + \text{acid} \rightleftharpoons \text{acid} + \text{base}$$

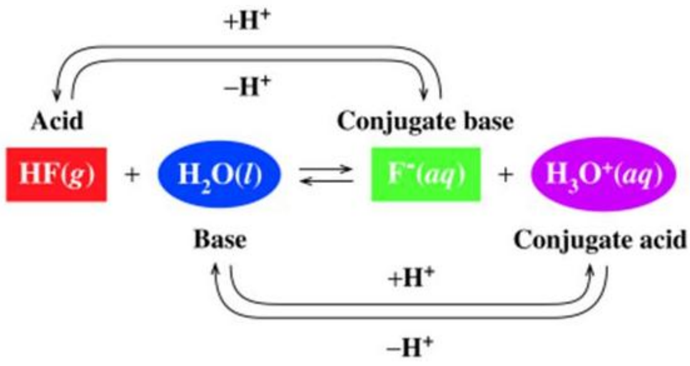
Got it: According to the Brønsted-Lowry theory, water can be an acid or a base depending on the substance reacting with it.

Water is a base when it is put with hydrochloric acid because water will gain a hydrogen ion to become H_3O^+. However, water will be an acid when it reacts with ammonia (NH_3) because water will lose a hydrogen ion (H^+) and become the base hydroxide ion, (OH^-).

Conjugate acid-base pairs

Conjugate acid-base pairs are

An example is shown below:



HF and F^- form one acid-base pair.

H_2O and H_3O^- form a second conjugate acid-base pair.

Each pair is related by the loss and gain of H^+

Unit 2.6.4 Strong and weak acids and bases

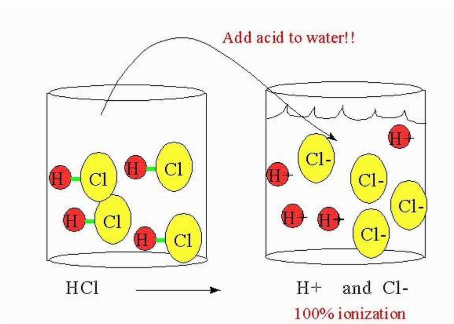
Acids

Strong acids are those that ionize completely in solution and therefore release all the H^+ possible. Examples of strong acids are hydrochloric acid, sulfuric acid and nitric acid which
ionize as follows:

$$HCl \rightarrow H^+ + Cl^-$$

$$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$$

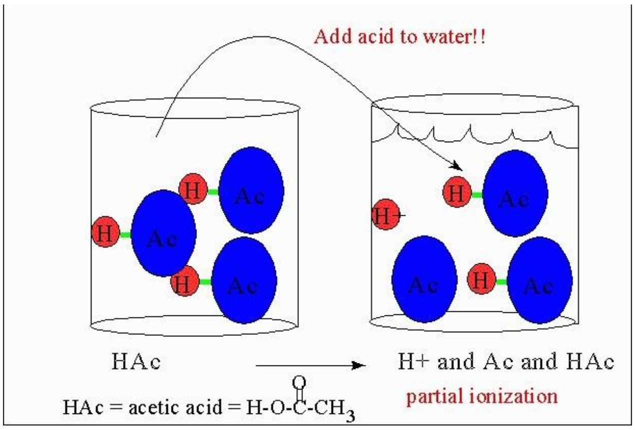
$$HNO_3 \rightarrow H^+ + NO^{3-}$$



Weak acids ionize partially in solution. Instead, they form a n equilibrium reaction. As a result not all of the possible H^+ ions are released into the solution. Strong acids have a lower pH (are more acidic) than weak acids even when they are at the same concentration.

An example of a weak acid is ethanoic acid (acetic acid or vinegar) as follows:

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$



Acid strength and acid concentration

The strength of an acid is a measure of the degree of its ionization. Strong acids are fully ionised but weak acids are only partly ionised. An acid that is not strong is called a weak acid.

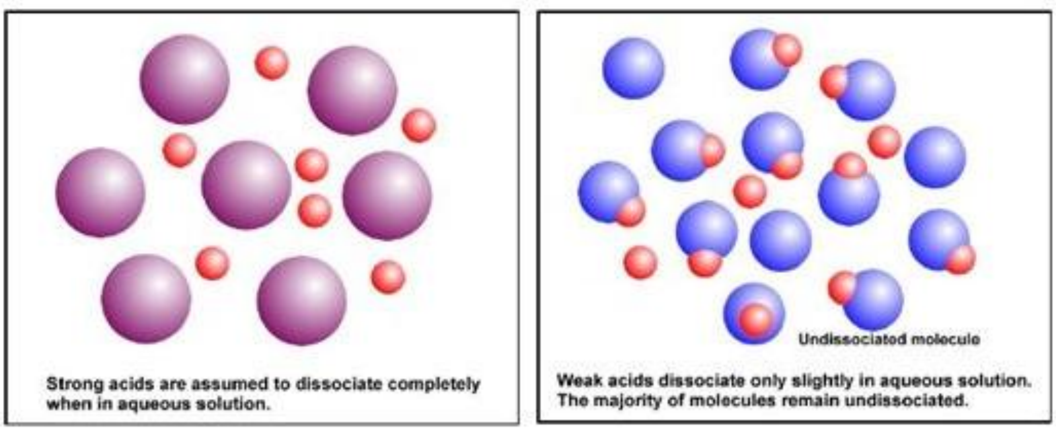
The concentration of an acid is a measure of the number of moles of acid in 1 dm^3 of acid solution. For example, 2 $mol.dm^3$ hydrochloric acid is twice as concentrated as 1 $mol.dm^3$ hydrochloric acid or 1 $mol.dm^3$ ethanoic acid. An acid that is not concentrated is called a dilute acid.

Base strength and base concentration

A **strong base** is one that ionizes completely in solution. The reaction is not reversible and the total number of OH^- ions are released. Some examples of strong bases are sodium hydroxide and potassium hydroxide:

$$NaOH(s) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq)$$

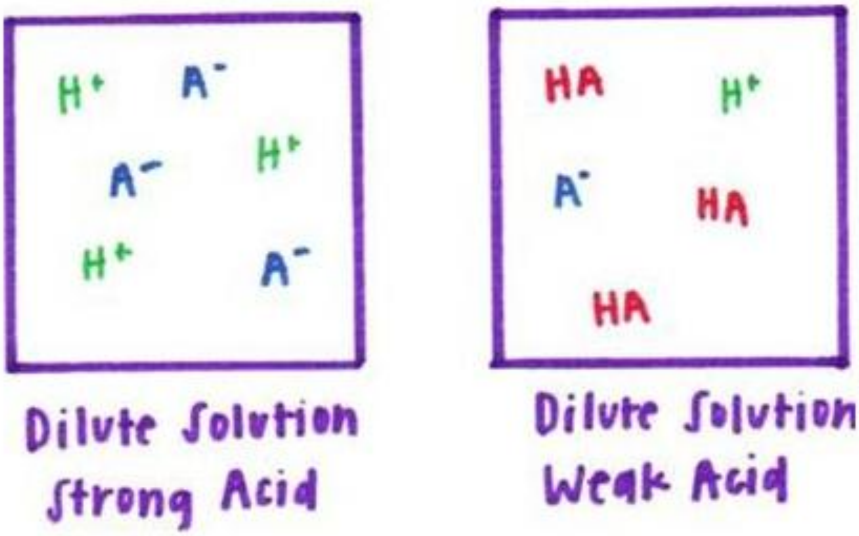
$$KOH(s) + H_2O(l) \rightarrow K^+(aq) + OH^-(aq)$$



A **weak base** ionizes only partially in solution. It is a reversible reaction and an equilibrium is established. An example of a weak base is an ammonia solution:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

The equilibrium favours the left hand reaction so only a few hydroxyl ions (OH^-) form, resulting in a weak base.



The concentration of an base is a measure of the number of moles of base in 1 dm^3 of base solution. For example, 2 $mol.dm^3$ sodium hydroxide solution is twice as concentrated as 1 $mol.dm^3$ sodium hydroxide solution or ammonia solution. A base that is not concentrated is called a dilute base.

Got it: **Strong** and **weak** refer to the ability of an acid or base to ionise. **Concentrated** and **dilute** refer to the quantity of acid or base present in the aqueous solution.

Unit 2.6.5 Chemical indicators

A chemical indicator is a substance that gives a visible sign, usually by a **colour change**, of the presence or absence an acid or an alkali in a solution. (Note: indicators can be used for other types of reactions, but for your purposes only neutralization reactions are being considered).

You have already met a simple indicator, litmus paper, which is red in acidic conditions and blue in alkaline conditions.

Here are a few other indicators that are quite common in a chemistry laboratory:

Indicator	Acidic	Neutral	Alkaline
Methyl orange	red	yellow	yellow
Phenolphthalein	colourless	colourless	pink

A universal indicator is a mixture of a different indicators and can be used to measure the approximate pH of a solution across the whole pH range from 1-14. The range of colours is similar to the rainbow, with reds being very acidic, greens being closer to neutral and the purples being very basic. Universal indicators give estimates of the pH – the values are not very accurate.

Did you know? Several plants, fruits and vegetables can act as chemical indicators. For example:

- Beetroot: A very basic solution (high pH) will change the color of beets or beet juice from red to purple.
- Curry powder: Curry contains the pigment curcumin, which changes from yellow at pH 7.4 to red at pH 8.6 Turmeric is a source of curcumin.
- Onions: Onions are olfactory indicators, meaning detected by smell. In acid solutions they smell strongly but you don't smell onions in strongly basic solutions. Red onion also changes from pale red in an acidic solution to green in a basic solution.
- Red cabbage: Red cabbage contains a number of different pigments (colour-producing chemicals) which detect a wide range of different pH values.



Unit 2.6.6 Some common types of reactions with acids

1. Acids will react with metal oxides, such as copper oxide, to make a salt and water.
   The general equation is as follows:
   acid + metal oxide → salt + water

   For example the reaction between copper oxide and hydrochloric acid:

   $$CuO(s) + HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$$
   metal oxide + acid → salt + water

2. Acids will react with metal hydroxides, such as calcium hydroxide, to make a salt and water.
   The general equation is as follows:
   acid + metal hydroxide → salt + water

   For example the reaction between calcium hydroxide and hydrochloric acid:

   $$Ca(OH)_2(s) + 2HCl\ (aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

   metal hydroxide + acid → salt + water

3. Acids will react with metal carbonates, such as, to make a salt, water and carbon dioxide.
   The general equation is as follows:

acid + metal carbonate → salt + water + carbon dioxide

For example, the reaction between calcium carbonate (found in chalk, limestone and marble) and hydrochloric acid:

$$CaCO_3(aq) + HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

metal carbonate + acid → salt + carbon dioxide + water

Again you would see bubbling during this reaction. Can you see why? Yes, it is because of carbon dioxide gas being released. This gas can be identified as CO_2 by passing it through a solution of limewater and seeing the formation of a white solid (called a precipitate).

The damage caused to rocks and buildings by acid rain is due to this reaction.

4. Acids will react with reactive metals, such as magnesium and zinc, to make a salt and hydrogen.
   The general equation is as follows:

acid + metal → salt + hydrogen

For example, the reaction between zinc and hydrochloric acid:

$$Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

metal + acid → salt + hydrogen

If you did this reaction you would see it bubbling. This is because the hydrogen is released as a gas. A simple test for hydrogen gas is that it stops a lighted piece of wood (called a splint in the laboratory) from burning.

Activity 2.2.20:

Answer the following multiple choice questions:

1. Which of these acids is least likely to be dangerous?
    a. Nitric acid
    b. Acetic acid
    c. Hydrochloric acid

2. Which statement about bases is true?
    a. All bases are soluble
    b. Bases can neutralize acids
    c. All bases are alkalis
3. Which statement about alkalis is true
    a. All alkalis are not bases
    b. All alkalis are soluble
    c. Alkalis cannot neutralize acids
4. Litmus paper that is placed in an acidic solution
    a. Turns red litmus paper blue
    b. Turns yellow litmus paper green
    c. Turns blue litmus paper red
5. A liquid is tested for pH using a universal indicator. The colour shown is green. This means that the liquid is
    a. neutral
    b. acidic
    c. basic
6. A liquid that has a pH of 6.5 is
    a. Neutral
    b. Weakly acidic
    c. Weakly basic
7. The contents of your stomach have a pH of 1. What does this mean?
    a. The stomach is a strongly alkaline environment
    b. The stomach is a neutral environment
    c. The stomach is a strongly acidic environment.
8. When an acid reacts with a metal hydroxide the products formed are:
    a. A salt and water
    b. A salt and hydrogen
    c. A salt, carbon dioxide and water

9. The products of a reaction with acid are A salt, carbon dioxide and water. What is the substance reacting with the acid?
    a. A metal oxide
    b. A metal carbonate
    c. A metal
10. Which acid could be used to manufacture a fertilizer called ammonium nitrate?
    a. Sulfuric acid
    b. Citric acid
    c. Nitric acid
11. What gas is produced when magnesium reacts with sulfuric acid?
    a. Oxygen
    b. Carbon dioxide
    c. Hydrogen
12. What salt is formed when copper oxide reacts with nitric acid?
    a. Copper sulfate
    b. Copper nitrate
    c. Copper chloride
13. A Brønsted-Lowry base is
    a. A proton donor
    b. A proton acceptor
    c. An electron acceptor
14. Which of the following statements is true?
    a. If you mix an acid with a base the base becomes stronger
    b. If you mix an acid with a base they neutralize each other
    c. If you mix an acid with a base no reaction takes place
15. A hydrogen ion (H^+) is the same as a
    a. Proton
    b. Neutron
    c. Electron

Answers 2.17

COMPONENT 1[NAME of sub-field e.g. physics, history etc.]

Introduction

[*Please write a brief overview of the component.Keep this as brief as possible, and word it so that it grabs the interest of the student, and encourages them to want to explore the theme further.*]

[Component 1]Content Structure

Topic Heading	Sub-Topic (with Approximate Instructional Time)
Topic heading	1. Sub-Topic (X hours) 2. Sub-Topic (X hours) 3. Sub-Topic (X hours)
Topic heading	4. Sub-Topic (X hours) 5. Sub-Topic (X hours) 6. Sub-Topic (X hours)
Topic heading	7. Sub-Topic (X hours) 8. Sub-Topic (X hours) 9. Sub-Topic (X hours)

Topic 1. [Topic Name]

Introduction

[Please write an introduction to this section that hooks the student's interest, and gives them an idea of the key learning involved in this section.]

Sub-topic 1. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

[Present content using short sentences and paragraphs, with illustrations wherever possible to keep the text from appearing too dense, and to facilitate understanding of concepts.]

[***Under each Unit please include the following features***:]

Highlight difficult **terminology** using bold font, and provide a clear explanation in a word box. Repeat these terms and explanations in the glossary. An example word box is shown alongside.

Terminology: explanation of the term in simple language

EXAMPLE: Where you are describing an example of something, use indented Arial font, 11 point, to separate it from the other text.

MAIN IDEA:{Give the key idea after a few paragraphs to help learners to identify the main concepts}

Activity X: Have you understood your reading:

Guidelines for the activity

Expected outcome of activity

Key learning from the activity

Assessment X: Description

Questions for formative self-assessment

(*Solutions to assessment questions to be included in the "Solutions" section, at the end of the workbook*).

1.2. Heading

1.3. Heading

My Notes

Use this space to write your own questions, comments or key points.

- Leave this blank for students to fill in their own comments

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

{After all of the units of a sub-topic include the following:

Summary of key learning:

- At the end of each sub-topic please give a brief summary of the key concepts / skills that have been covered in bullet points.
-

Summary Assessment

Questions for assessment of this sub-topic

(*Solutions to assessment questions to be included in the "Solutions" section, at the end of the workbook*).

Suggested sources of additional information

[*Please include here the links to helpful resources such as websites, Youtube videos, etc. As indicated above, as much as possible **open source material** should be utilised and referenced.*]

Sub-topic 2. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 3. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

etc.

[*Continue in this manner, with as many sub-topics and units as needed for this topic.*]

Topic 2. [Topic Name]

Introduction

[Please write an introduction to this section that hooks the student's interest, and gives them an idea of the key learning involved in this section.]

Sub-topic 1. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 2. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

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-

1.1. Heading

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1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 3. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

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-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

etc.

[*Continue in this manner, with as many sub-topics and units as needed for this topic.*]

Topic 3. [Topic Name]

Introduction

[Please write an introduction to this section that hooks the student's interest, and gives them an idea of the key learning involved in this section.]

Sub-topic 1. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 2. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 3. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

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-

1.1. Heading

1.2. Heading

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Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

etc.

[*Continue in this manner, with as many sub-topics and units as needed for this topic.*]

COMPONENT 2[NAME of sub-field e.g. physics, history etc.]

Introduction

[*Please write a brief overview of the component.Keep this as brief as possible, and word it so that it grabs the interest of the student, and encourages them to want to explore the theme further.*]

[Component 2]Content Structure

Topic Heading	Sub-Topic (with Approximate Instructional Time)
Topic heading	10. Sub-Topic (X hours) 11. Sub-Topic (X hours) 12. Sub-Topic (X hours)
Topic heading	13. Sub-Topic (X hours) 14. Sub-Topic (X hours) 15. Sub-Topic (X hours)
Topic heading	16. Sub-Topic (X hours) 17. Sub-Topic (X hours) 18. Sub-Topic (X hours)

Topic 1. [Topic Name]

Introduction

[Please write an introduction to this section that hooks the student's interest, and gives them an idea of the key learning involved in this section.]

Sub-topic 1. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 2. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 3. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

etc.

[*Continue in this manner, with as many sub-topics and units as needed for this topic.*]

Topic 2. [Topic Name]

Introduction

[Please write an introduction to this section that hooks the student's interest, and gives them an idea of the key learning involved in this section.]

Sub-topic 1. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 2. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

Sub-topic 3. [Topic name]

Content:

[Paste the content outline in the form of Units:

Unit 1: [Name of unit]

Unit 2: [Name of Unit]

Unit 1. [Name of unit]

Learning outcomes:

When you have completed this unit, you should be able to:

- *Write the list of learning outcomes as they are in the approved NASCA curriculum that link with this unit (they don't have to be in the order in which they appear in the curriculum)*
-

1.1. Heading

1.2. Heading

1.3. Heading

Unit 2. [Name of unit]

2.1. Heading

2.2. Heading

2.3. Heading

Unit 3. [Name of unit]

3.1. Heading

3.2. Heading

3.3. Heading

etc.

[*Continue in this manner, with as many sub-topics and units as needed for this topic.*]

etc.

SOLUTIONS

[*Please provide the solutions to the Activities where relevant, and to the Assessment exercises within the workbook.****Please make sure that these are clearly and correctly labelled to link with the questions***.]

REFERENCES

[*Please provide a full list of references, using Harvard referencing style*.]

GlOSSARY OF TERMS

EXEMPLAR(S)