NASCA Chemistry Materials Draft 1

Units 7-10

DRAFT

Natural Science

Chemistry

Unit 7: The Mole concept

Outcomes

By the end of the unit, you should be able to:

- Define what is a chemical reaction
- Write simple chemical equations
- Describe one mole of a substance the unit of amount containing Avogadro's number of particles;
- Define molar mass;
- Calculate the molar mass from the chemical formulae of chemical compounds, using the periodic table.

Introduction to Unit 7

You have already learned how to write chemical formulae for compounds. In this unit, you will use these formulae for writing chemical equations. These chemical equations are very similar in many ways to the equations you write in maths, although there are obviously many differences.
The most important thing you will learn in this unit is a new unit for measuring amount of substance, called the mole. On the one hand, it is very easy to calculate but on the other, it is more difficult to understand what it is.

Chemical shorthand - a universal language!

When you are communicating with your friends by sms, you may have learnt a special language called sms language. You friend may send you a message "CU 2moro" (see you tomorrow) and you may reply "Gr8" (great) and everybody who understands sms language would be able to read it!
Here is a short warm up activity

Activity 7.1

Write the following sms phrases in English:

1. 10Q
2. TGTBT
3. G2GICYAL8ER

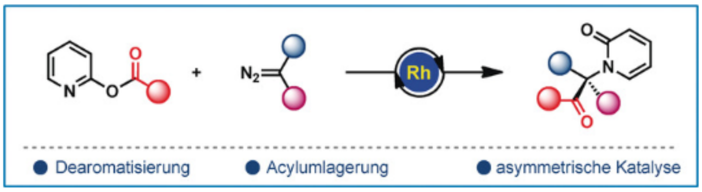
For those of you who are not used to the language, here are the translations

1. Thank you
2. too good to be true
3. Got to go I'll see you later

Chemists also have their own language, which they use to write down what happens in a chemical reaction. Unlike sms language, it is independent of what language a chemist speaks.
A German, English or Chinese person would all be able to understand the information given by a chemical equation written in chemical shorthand, even though they might not understand if they were told about the reaction in words. For example, figure 7.1 shows an extract from a German chemistry journal.

VIP: Access to N-Substituted 2-Pyridones by Catalytic Intermolecular Dearomatization and 1,4-Acyl Transfer

Guangyang Xu, Ping Chen, Pei Liu, Shengbiao Tang, Prof. Dr. Xinhao Zhang, Prof. Dr. Jiangtao Sun



Metall-Carben-Reaktion: Eine rhodiumkatalysierte Dearomatisierung überführt Pyridine mit O-Substituenten in N-substituierte 2-Pyridone. Rechnungen stützen als Mechanismus eine Pyridinium-Ylid-Bildung mit anschließender 1,4-Acyl-Verschiebung von O zu C. Eine asymmetrische Variante mit chiralen Dirhodiumkomplexen als Katalysatoren verläuft hoch enantioselektiv.

Fig 7.1 extract from a German chemistry journal
https://onlinelibrary.wiley.com/page/journal/15213757/homepage/news#_vchnews45680

Unless you understand German, you will not be able to make sense of the writing but if you know chemistry, you will recognise the chemical symbols and equations in the diagrams.

So you see that understanding chemical symbols and formulae is very useful for people to communicate with each other even if they speak different languages.

What is a chemical reaction?

Compare the situations at the start and at the end

When you mix substances, there are at least two possible outcomes:
1. A mixture will be formed. For example, if you mix iodine crystals with flour you will end up with the exact substances you started with (iodine and flour) except that they will now be in a bit of a mess, i.e. forming a mixture! If you really needed to, you could separate the mixture of iodine and flour again by sifting or with a small pair of tweezers.
2. The substances may react with each other. If you mixed iodine crystals with small pieces of sodium metal instead of flour, you would get quite a fright! Instead of just forming a mixture, the 2 substances react with each other quite violently, making noise and giving off *smoke*. When all the action is over you will find there is something completely new in the container and no iodine nor sodium to be seen. Click on the link below to see the reaction
https://www.youtube.com/watch?v=sxwrGwPu51g

In simple terms, this is what happens during a chemical reaction. You start with a substance or substances and end up with something new. The original substances, which are changed, are called reactants and those formed are called products of the reaction. It is common to find that the physical properties of reactants and products are completely different. In the example of iodine and sodium above you would find that dark crystals of iodine and a silverish metal called sodium produce a white solid, sodium iodide.
In order to represent reactions like the one between sodium and iodine, or any reaction, for that matter, at the symbolic level, you write a chemical equation. The first step in doing this is to write the correct formulae of the reactants on the left, separated by a plus sign, then an arrow followed by the formulae of the products on the right-hand side. An equation representing the reaction between the imaginary substances A and B is shown below:

$$A(g) + B(g) \rightarrow C(s) + D(\ell)$$

The plus sign means *reacts with*, and the arrow means *produces*. Usually one also shows the physical states of the reactants and products in an equation, i.e. whether each substance is a liquid (ℓ), a solid (s) or a gas (g). If you look at the equation above and you will notice that two gases react with each other to form a liquid product (D) and a solid product (C).

What you have written down on paper is called the *chemical equation*. What has happened in the test tube is the *chemical reaction*! A *chemical equation* is used to represent a *chemical reaction* - they are not the same! The one is actually done in the laboratory (the reaction) while the other is written on paper (the equation).

Sometimes reactants or products are dissolved in water. We say they are in an *aqueous solution* and this is represented by (aq) after the formulae, for example

$$HC\ell(aq) + NaOH(aq) \rightarrow NaC\ell(aq) + H_2O(\ell)$$

Sometimes it is very important to know the states of reactants and/or products and it is a good idea to get into the habit of always showing the states whenever you write a chemical equation. A knowledge of the physical states completes the picture of a reaction!

You will encounter some cases where you are given an equation in which the reactants and products are identical but differ only in their physical states for example:

$$H_2O(\ell) \rightarrow H_2O(s)$$

An equation like this does not represent a chemical reaction as no new products have been formed. This equation represents a phase change, from liquid to solid, in other words, freezing.

Chemical equations have to be balanced

You may remember in unit 3 we talked about Dalton and his laws. One of them was that matter cannot be created or destroyed. This law is also called the *law of conservation of matter*. Having written out your chemical equations with the correct formulae, you also need to check that you have not "lost" any atoms along the way. To see what we mean by this, consider the equation below, which represents the reaction between hydrogen gas and oxygen, gas to form water.

You would write

$$H_2(g) + O_2(g) \rightarrow H_2O(\ell) \text{ (equation 1)}$$

Chemistry can be compared to cooking! If you bake a cake and you use 2 eggs in the mixture, then when the cake is ready you know that there must be 2 eggs in the cake, even if you cannot see the eggs. Things cannot just disappear!

This is also true for chemical reactions. Count the number of H- atoms and O - atoms on each side of equation 1. You will see that there are 2 H- atoms on each side, but there is a problem with the O- atoms. There are 2 O- atoms in the reactants and only 1 O- atom in the products. According to the law of conservation of mass, if 2 atoms of oxygen are present in the reactants of equation 1, then there must be 2 atoms of oxygen in the products as well.

If you analyse what is happening in equation 1, it is clear that 2 H-atoms combine with only 1 O - atom to form H_2O. This means if we use ONLY ONE molecule of each of the gases H_2 and O_2 the equation between should be written:

1 molecule $H_2(g)$ + 1 molecule $O_2(g)$ + 1 molecule $H_2O(\ell)$ + 1 O-atom (left over)

There is still 1 O- atom available to react with another H_2 molecule to form a second molecule of H_2O. Therefore, the reaction actually happens between 2 molecules of H_2 and only 1 of O_2 and 2 molecules (not one) of H_2O are formed. So the equation looks like this:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) \text{ (equation 2)}$$

What you have done, is balance the equation by writing *2* before the H_2 and H_2O as shown above. Equation 2 is balanced because if you count all the atoms of each type in the reactants and products you will see that the totals are the same.

It is very easy to balance chemical equations. You do not have to go through the whole process of reasoning which you did for the reaction between $H_2(g)$ and $O_2(g)$! All you have to do is to count atoms on the left and on the right and then balance the numbers by writing 2's or 3's or 4's in front of the formulae. You may not alter the formula you have written unless you suspect it is incorrect!

For example, in the equation representing the reaction between magnesium metal and oxygen gas to form magnesium oxide, the unbalanced equation would be:

$$Mg(s) + O_2(g) \rightarrow MgO(s)$$

You MAY NOT just change the formula of magnesium oxide to balance the equation! Some students are tempted to do this: $Mg(s) + O_2(g) \rightarrow MgO_2(s)$. Although the equation is now balanced, the formula for magnesium oxide IS INCORRECT. Once you change the formulae of a compound, you change the compound. MgO_2 does not exist. If it did exist, it would be a different substance with different properties. The correct way to balance this equation is shown below:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Although we give you a strategy below, you must be aware that there is no foolproof way of balancing equations.

Strategy for balancing equations

Say for example that you want to write a balanced equation representing the reaction of methane. gas with oxygen gas to form carbon dioxide gas and water vapour (also a gas). This reaction is also called the combustion reaction of methane.

Follow the steps below:

Step 1:

Work out what the reactants and products are. In this case you are told clearly that methane and oxygen react. Therefore, these are the reactants. You are also told that the carbon dioxide and water are formed. These are the products.

Therefore the equation in words is: methane gas plus oxygen gas forms carbon dioxide gas and water vapour.

Step 2:

Write the formula for all the reactants and products:

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

Step 3:

Count the number of atoms of each different kind on the left and right side of the equation.

Carbon: There is 1 on the left (in CH_4) and 1 on the right (in CO_2). There are the same number C-atoms on either side of the equation, so the equation is balanced in terms of carbon!

Hydrogen: There are 4 on left and 2 on right side. This is a problem so you have to alter the equation so that there are the same number on each side. Therefore, multiply the molecule on the right side containing hydrogen atoms by 2 to make a total of 4 hydrogens. The equation now looks like this:

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

If you check by counting the number of atoms of each kind on both sides, you will see that the equation is not yet balanced!! We still have to check the oxygen atoms.

Oxygen: There are 2 atoms on the left side (from O_2) and 4 atoms on the right side (from CO_2 and $2H_2O$). This is a problem so you have to alter the equation so that there are the same number on each side. Therefore, multiply the oxygen molecule on the left side by 2 to make a total of 4 oxygen atoms. Your equation now looks like this:

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

You can check again, but now the equation is balanced. This means that the total of the atoms on the left hand side is the same as on the right hand side.

You will need to balance your chemical equations every time you write them, as this is important in calculations.

From the above example, you can see that there are 3 things you need to get right in order to end up with a balanced equation.

1. You need to know what all the reactants and products are.
2. You need to write the correct formulae of the reactants and products and their states (gas, liquid or solid)
3. Finally you need to go through the process of balancing the equation without changing the formulae.

The following simulation will help you with step 3 and will allow you to become familiar with some formulae.

Activity 7.2

This activity requires access to a PHET simulation

Go to the following web address:

https://phet.colorado.edu/en/simulation/balancing-chemical-equations

Click on the Play arrow in *Balancing Chemical Equations*: Introduction. You will see a screen as shown in figure 7.2

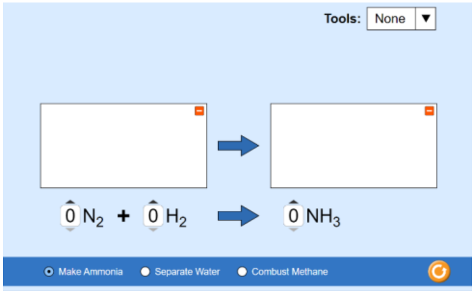


Fig 7.2 opening screen of PHET simulation- balancing chemical equations

Select *combust methane* and select the  tool and you will see the screen below (figure 7.3)

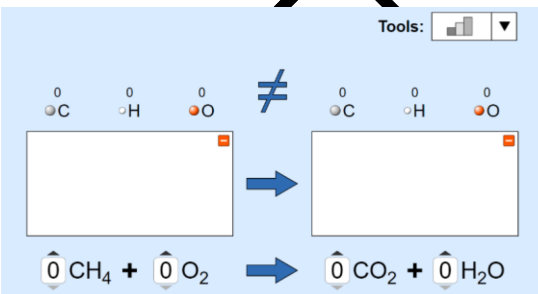


Fig 7.3 screen for balancing *combust methane*

Use the up and down arrows to adjust the number for each molecule.

Remember, this is the same equation we have just balanced above.

When you finally have the correct equation, you will see the screen below in figure 7.4:

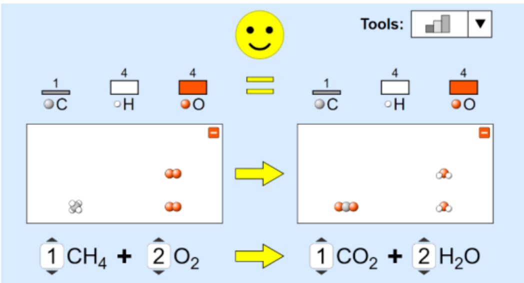


Fig 7.4 correct answer screen for balancing *combust methane*

Now try the other two activities, *make ammonia* and *separate water*

Guided reflection

Notice that for each equation in the simulation, you are told the formulae for all the reactants and products and you are not asked to mention the states. In fact, in the two examples, all the substances are gases. Below we write the correct equations with the states included. I hope you got the correct answers.

For *make ammonia* the correct full balanced equation is $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

This equation is very important as it shows the reaction for making ammonia gas from the nitrogen in the air that is needed for making fertilisers for food production.

For *separate water* the correct full equation is $2H_2O(g) \rightarrow 2H_2(g) + O_2(g)$

This equation shows the reaction showing how water can be broken down into hydrogen and oxygen.

If you now press the link at the bottom of the screen that says *game* you can get more experience in balancing equations. See if you can get a perfect score for level 3!!

Final guidelines for writing chemical equations

Before you practice writing and balancing your own equations, it will be helpful to pull everything together just to remind you of all the things you have to bear in mind. Writing chemical equations is easy if you stick to a few rules. Most of the things you should remember are listed below. You find an opportunity to practice a little later on.

1. You must write the correct chemical formula for every element or compound that is going to appear in your equation. What this means is that there is *only one* correct formula for each element or compound - you cannot change this to suit yourself or to balance an equation. Remember you can only balance an equation by writing numbers in front of a formula, NEVER by changing a formula. For example, sodium oxide is always Na_2O and *never* NaO. You have learned to write these formulae so always be careful and get them right.
2. You must know which elements exist as diatomic molecules. It is easy to remember them. Go to your periodic table and look for a pattern to help you remember — you should know that hydrogen gas is H_2! Then if you go down to the 2nd period to nitrogen (N_2) then across to oxygen (O_2) and fluorine (F_2) then change direction and go down the group of halogens Cl_2, Br_2 and I_2 . So the common gases and halogens are written as diatomic molecules.
3. All other elements are represented as being monatomic unless you are given information about them. For example, you can represent sulfur as S unless you are told that it is S_8. You will not know this unless you are told!
4. Remember that combustion means *reacts with oxygen*!
5. Some compounds have special names. You should know most of them, but just to remind you, a few are given below. Because most acids are already dissolved in water, they are represented as aqueous solutions:
   sulfuric acid $H_2SO_4(aq)$
   hydrochloric acid HCl(aq)
   ammonia $NH_3(g)$
   methane $CH_4(g)$
6. The final step is to balance the equation. You do this by counting atoms on the left hand side and right hand side of the equation and writing numbers in front of a formula to balance the numbers. If your equation seems to be impossible to balance, you should suspect that you have written one of your chemical formulae incorrectly. One way of checking the correct formula is to go to google and asking for the chemical formula but be careful because google is not accurate in putting the subscripts (the numbers that we write below the symbol).

Activity 7.3 will give you some practice in writing your own equations.

Activity 7.3

Try writing balanced equations for the following reactions. It is not only easy, but is satisfying and fun to do once you get the hang of it.

1. The combustion of magnesium ribbon to form a white powder, which is magnesium oxide.
2. Carbon burns in oxygen in the air to form carbon dioxide.
3. Hydrochloric acid reacts with sodium carbonate to form sodium chloride, carbon dioxide and water.
4. Glucose powder ($C_6H_{12}O_6$) burns in oxygen to form carbon dioxide and water vapour. This is also a combustion reaction.

Guided Reflection

1. $2Mg\ (s) + O_2(g) \rightarrow 2MgO\ (s)$
2. $C(s) + O_2(g) \rightarrow CO_2\ (g)$
3. $HCl\ (aq) + Na_2CO_3 \rightarrow 2NaCl\ (aq) + CO_2(g) + H_2O\ (\ell)$
4. $C_6H_{12}O_6\ (s) + O_2(g) \rightarrow 6\ CO_2(g) + 6H_2O\ (g)$

Now you have learned how to write and balance equations. In the next section you will be introduced to the "mole" which is also known as the chemist's dozen.

The mole as a counting number

Sometimes we count in two's or twelve's!

We all learn to count at an early stage. First, we learn to count one, two, three etc. and then in two's: two, four, six, eight. The reason for using different counting numbers is that this can make counting things faster and easier.

The counting unit used depends on what you are counting. You count your shoes in pairs because most people wear 2 shoes and you cannot buy only 1 shoe in shops!

In the supermarket, eggs are packaged in packs of 12 called a dozen or in packs of 6 known as a half-dozen. You no longer have to count the individual eggs every time to know how many eggs you are buying but count the packs (dozens).

As you know atoms are very small particles and if you wanted to count them in pairs or dozens this would take a very long time. Therefore, chemists have found a counting number, which enables us to count groups of atoms as a unit. This number is called a *mole*.

You will later see why this particular counting number is important and why it became the unit for *amount* of substance in chemistry.

So when you count shoes you would count one pair, two pairs, three pairs, etc. and because you know how many items there are in a pair you would know that if you have 8 pairs then you actually have 16 shoes.

When you count eggs, you count 1 dozen, 2 dozen etc. If you buy 8 dozen eggs then you know that you have $8 \times 12 = 96$ eggs.

Similarly, when you count in moles you count 1 mole, 2 moles, 3 moles, etc. and all you have to know is how many items there are in 1 mole so that you can work out how many you have.

The number which is in a pair is 2

The number which is in a dozen is 12

The number which is in a mole is 602200000000000000000000 or 6.022×10^{23} (which is easier to write!)

Why this number? Why is it so big? Where does it come from?

In the last section, we introduced you to a large number without giving any reason. In this section, we will try to show you where this number comes from and why we need it.

It is easy to see why some food items like eggs and bread rolls are counted in dozens. Shoppers usually need more than one, and often they need a number as large as 12.

In the case of eggs and bread rolls we can see them and count them
You saw in the previous section that when substances react with one another, they react in whole numbers of atoms to make a unique formula.
For example, in the second question in activity 7.3, the equation for the reaction of carbon with oxygen was
$C(s) + O_2(g) \rightarrow CO_2\ (g)$
One atom of carbon reacts with one molecule of oxygen to give one molecule of carbon dioxide.
The problem is that atoms and molecules are two small to see, so how can we count the right number of atoms and molecules to be sure that they match?
We can count eggs but we can't count atoms. They are too small.
Well, one example that can help us work this out is to think of how tellers count money in the bank. Let us suppose that Mr Mthethwa runs a spaza shop and in a week, he collects a large number of 10c coins. At the end of the week, he goes to deposit his money in the bank. Before he goes to the bank, he used small plastic bags from the bank with a zip opening as in figure 7.5

Figure 7.5: Bank bag with coins
https://media.istockphoto.com/photos/plastic-bank-bag-filled-with-south-african-cents-this-image-can-be-picture-id1029814936?s=2048x2048

If the bag contains only 10c coins, how can the teller in the bank find out how much money is in the bag? Try activity 7.4

Activity 7.4
In the bank they have machines that weigh the bags. Each 10c coin weighs 2g. Mr Mthethwa arrives at the bank with several bags with the total amount of money written on the bag.
How many coins should be in the bag?
How many coins are in the bag?
Has Mr Mthethwa made any mistakes? Fill in the table below.

Amount of money written on the bag	Number of coins that should be in the bag	Mass of the bag	Number of coins actually in the bag	Correct amount of money
R50	500	1kg (1000g)		
R25		450g		
R40		400g		
R35		750g		

Guided reflection
Here is the table completed. You can see that Mr Mthethwa has made some mistakes. He will have to go back and count the coins again.

Amount of money written on the bag	Number of coins that should be in the bag	Mass of the bag	Number of coins actually in the bag	Correct amount of money
R50	500	1kg (1000g)	500	R50
R25	250	450g	225	R22.50
R40	400	400g	200	R40
R35	350	750g	375	R37.50

This exercise introduces us to an interesting idea. Each coin has a certain value that is directly connected to its mass. We call this the *amount* of money. The bigger coins are worth a larger amount of money.

In the case of atoms, we have learnt that different atoms have different masses just like coins only the atoms are so small that we can't see them. But it works out that if we take 6.022×10^{23} atoms of any atom, we will get the mass of the atoms in grams.

Look at the following examples:

- The atomic mass of hydrogen is 1 atomic mass unit (amu) and if we take 6.022×10^{23} atoms of hydrogen, we will have 1g
- The atomic mass of carbon is 12 atomic mass units (amu) and if we take 6.022×10^{23} atoms of carbon, we will have 12g
- The atomic mass of oxygen is 16 atomic mass units (amu) and if we take 6.022×10^{23} atoms of oxygen, we will have 16g

Looking at the equation for the reaction of carbon with oxygen

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

One atom of C will have a mass of 12amu

One molecule of oxygen (O_2) will have a mass of 2 x 16 = 32 amu (Notice there are 2 O atoms to make the molecule so we have to multiply by 2).

So 12 amu of carbon will react with 32 amu of oxygen. If we take 6.022×10^{23} of each, then 12 g of carbon will react with 32 g of oxygen

The mole is defined as the official unit for amount of substance and taking 1 mol of substance gives us 6.022×10^{23} atoms and it gives us the mass of the atom in grams.

The mass of the atom in grams is called the molar mass.

So the molar mass of carbon is 12g and mass of a single carbon atom is 12 amu.

This means that there are 6.022×10^{23} amu in 1 gram.

Therefore when you want to know the molar mass (or the mass of one mole), you look it up on the periodic table and express it in grams.

I thought a mole was a small animal!

Figure 7.6: a mole

https://encrypted-tbn0.gstatic.com/images?q=tbn:ANd9GcR9XpaRBQpctF-QcrbtVKYpbHWq9XQYEHMHCmiVUMOzOV-TH-Gm

If someone you know has a vegetable garden you may have heard them shouting when they find that some of their carrots have been eaten by a small animal which makes tunnels underground.

In English, the word “mole” can also refer to a large lump, which some people have on their faces or other parts of their bodies. Even the former president of the USA, Mr Barack Obama has one!!!

Figure 7.7: Mr Obama’s mole
https://www.scoopwhoop.com/what-does-the-mole-say/#.a26hzrbd6

But in chemistry the word comes from the word *moles* which means large mass, so don’t get confused about the meaning of the word. Remember the mole as the chemist’s dozen.

A mole of beans and a mole of atoms? What’s the connection?

Because it is very difficult to imagine what a mole of beans looks like, it is important that you complete activity 7.5. This activity will help you to compare the mass and volume of heaps of different substances, which each contain the same

Activity 7.5

Go and find 10 grains of rice, 10 sunflower seeds and 10 nuts. You can use any three sets of similar objects that are the same size and shape, such as beans or marbles, but they must be different from each other. Put the rice, sunflower seeds and nuts into separate heaps as shown in figure 7.8.

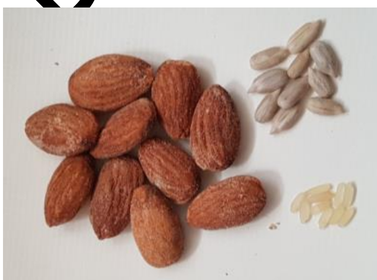


Figure 7.8: 10 nuts, 10 rice grains and 10 sunflower seeds
(own photo)

a. Do you think the masses of the 3 piles are the same? Why do you say so?
b. Is the volume of each pile the same? Explain.
c. What is the same about the 3 piles?

Guided Reflection

a. No! A grain of rice has a much smaller mass than a single nut or a sunflower seed, so it is clear that 10 grains of rice would have a smaller mass than 10 nuts or 10 sunflower seeds.

b. You can see just by looking at the three piles that the almonds has the greatest volume.

c. The only thing that is the same about the three piles is that there are 10 items in each pile. This is a very important idea when we start thinking about atoms because as you saw in the chemical equations above, atoms react with each other in fixed ratios.

From oranges to atoms!

Remember that you are studying chemistry so the rice, beans and stones in activity 7.5 can be compared to atoms of different sizes. The rice could represent some small atoms (say S), the beans could represent bigger atoms (say Al) while the stones could be big atoms (like Mg).

Now if the mole is a number used for counting things, this means that if you had 1 mole of eggs you would have 602200000000000000000000 eggs. If you had 1 mole of oranges you would have 602200000000000000000000 oranges. You have to try to use your imagination a bit now. Try to imagine a heap of 1 mole of eggs and another heap next to it of 1 mole of oranges. What would be different about the 2 heaps? What would be the same?

Activity 7.6 takes you from oranges to atoms! See if you can transfer what you know about objects you can see, to atoms, which you can't see.

Activity 7.6

Examine the pictures of 3 sets of test tubes in Figure 7.9 below.

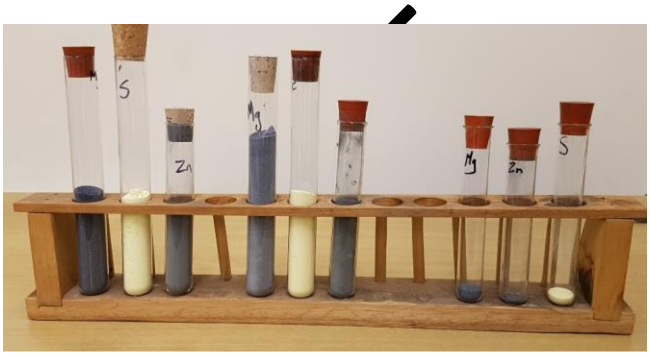


Figure 7.9: Sets of test tubes containing zinc, magnesium and sulfur
(own photo)

Set A: contains <u>equal volumes</u> of zinc, magnesium and sulfur (on the left).
Set B: contains <u>equal masses</u> of zinc, magnesium and sulfur (in the middle).
Set C: contains different volumes and masses of zinc, magnesium and sulfur (on the right).
One of the sets contains exactly one mole of each of the three elements. Which set is it most likely to be?
Explain your answer. You may go back to your beans and rice to assist in your explanation. You must also explain why the other sets do not contain molar quantities of zinc, magnesium and sulphur.

Guided Reflection

Set C is most likely to contain 6.022×10^{23} atoms of each element. Here is the reasoning:
Atoms of different elements do not have the same mass (you learnt this in unit x). So 1 S atom will not have the same mass as 1 Zn atom or 1 Mg atom. Therefore, if you have the same number of atoms of each in 3 separate test tubes, the masses will definitely be different. Therefore set A will not contain one mole of each element.
The same applies to volume. Atoms of different elements do not have the same volume – some are bigger than others are. So 1 S atom will not have the same volume as 1 Zn atom or 1 Mg atom.

Therefore, if you have the same number of atoms of each in 3 separate test tubes, the volumes will definitely be different. Therefore, set B will not contain one mole of each element.
So the set of test tubes with different masses and volume of each is most likely to contain one mole of each.

By now you should have a reasonably clear picture of how many atoms (or any item) you would have to count to have one mole.
Do you know how long it would take to count 1 mole? If you can count 60 beans in 1 minute (60 seconds), try to work out whether you would be able 1 mole of beans in your lifetime. You should get an answer of close to 19000000000000000 (or 1.9×10^{16}) years

Not only atoms are counted in moles!

Although an extremely large unit like the mole is suitable for counting atoms and molecules activity 7.7 makes the point that other things can also be counted in moles.

Activity 7.7

It has been found that the volume of 20 drops of water is 1 cm^3. The estimated volume of all the oceans in the world is thought to be 10^{16} m^3.

a. How many drops of water are there in the oceans? (Hint: First, convert m^3 to cm^3.)

b. What is the amount of drops in all oceans?

Guided Reflection

Volume of 1 drop of water = 1 cm^3 / 20 = 0.05 cm^3

a. Total volume of all the oceans = 10^{16} m^3 = $10^{16} \times (10^2)^3$ = 1×10^{22} cm^3.

There are 100 (10^2) cm in 1m. Then we have to raise it to the power of three for cubic metres.
Therefore number of drops in the all the oceans is

$$1 \times 10^{22} \text{ cm}^3 / 0.05 \text{ cm}^3 = 2 \times 10^{23} \text{ drops}$$

b. For the amount of drops, we use the unit moles. 6×10^{23} drops would correspond to 1 mole.

$$2 \times 10^{23} \text{ drops would be } 1/3 \text{ mole } (2 \times 10^{23} / 6 \times 10^{23})$$

This is an amazing answer. It is another way of showing how large the number corresponding to a mole is.

By now, you should be convinced that it is physically impossible to count a mole. In the next section, you will take a look at how you can supply somebody with 1 mole of S atoms even if you cannot count them.

The mole takes too long to count

6.022×10^{23}, why choose such a large number if one can't count it?

This number was chosen, based on calculations and experiments and to make chemists' lives easier. It is much easier working with a number that is related directly to the relative atomic mass of an element as it is given on the periodic table than with some other number and this is where the mole comes in. If you take the relative atomic mass of an element and express it in grams instead of amu, then you can show that it will contain the same number of atoms as the relative atomic mass of any other element (in grams).
The relative atomic mass of C is 12 amu. If you change the unit to grams instead of amu this is 12 g. If you look at S, which has a relative atomic mass of 32 amu, then what this means is that 12 g of C contains exactly the same number of atoms as 32 g of S. The actual number of atoms in 12g of C and 32g of S would be 6.022×10^{23} which is an amount of 1 mole (also abbreviated as 1 mol). This is very convenient, because to find how many grams of an element would contain an amount of 1 mol, you can go straight to the periodic table and use the relative atomic mass.

This number (6.022×10^{23}) is also known as Avogadro's number. Amedeo Avogadro was an early chemist who succeeded in persuading people of the time that equal volumes of gases at the same temperature and pressure contain the same number of molecules or atoms.

Figure 7.10: Amedeo Avogadro (1776 -1856)
https://www.sapaviva.com/amedeo-avogadro/

You may ask: 1 mol of Al does not have the same volume as 1 mol of Mg. How is it possible that 1 mol of CO_2 can occupy the same volume as 1 mol of He? *This is only true for gases*. This is because the size of individual atoms in a gas is so small compared to the volume occupied by the gas. The volume of 1 mol of gas at 0 °C and 101.3 kPa pressure is 22.4 litres. This is the molar gas volume (the volume of 1 mol of any gas).
From activity 7.4, you learnt that we could find out the *amount* of money in a bag of coins by weighing a bag with a known number of coins. You are able to do this because you know the mass of a single coin. The same is true of atoms. We know the mass of a single atom so we can find the amount (in this case of a very large number of coins) by weighing them.
If I ask you for 1 mol of Mg then I am actually asking for approximately 6.022×10^{23} atoms of Mg. I say approximately because we only know the number the third decimal place and this could mean millions of atoms.

Using moles in calculations

You must be able to use everything you have learned about the mole in calculations. This is one of the most important aspects of chemistry and is easy to do with a little bit of practice. To show you how to approach problems involving the mole as a counting number, look at the activity that follows. It will help you understand how to do problems involving the mole:

Activity 7.8

To focus your thoughts, think about weighing 64.132 g of sulfur (which is a yellow powder). How many moles of sulfur and how many atoms of sulfur would you have weighed?
In the explanation that follows, you must be aware that the words *mole* and *mol* are used as synonyms. They mean the same.

1. First, look up the relative atomic mass of sulfur: 32.07 amu according to your periodic table. It may be given to more significant figures on other tables. This means that 1 mol of sulfur has a mass of 32.07 g. A better way of saying it is that sulfur has a molar mass of 32.07 g/mol or g mol^{-1}. We say it as *gram per mol*.
2. Next, you need to find out the number of moles in 64.132 g.
   64,132g/32,066g mol^{-1} = 2.000 mol of sulfur. (If you look at the units, you will see that they cancel to give the unit of the answer, which is mol).
3. Finally, to find the number of atoms you have to go back to what you know! You know that 1 mol is 6.022×10^{23}
   So the number of atoms. = 6.022×10^{23} atoms/mol x 2.000 moles = 1.204×10^{24} atoms

4. Now follow the same steps to see if you can find out how many moles and how many atoms you have weighed if you weigh 68.97 g of sodium.

Guided reflection

Notice that this problem is about moving between macro representations (grams and mol) to sub-micro representations (amu and molecules).

1. First, look up the relative atomic mass of sodium: 22.99 amu according to your periodic table. This means that 1 mol of sodium has a mass of 22.99 g. In other words sodium has a molar mass of 22.99 g/mol or g mol^{-1}.
2. Next, you need to find out the number of moles in 68.97 g.
   68.97g/22.99g mol^{-1} = 3.000 mol of sulfur.
3. Finally, to find the number of atoms you know that 1 mol is 6.022×10^{23}
   So the number of atoms. = 6.022×10^{23} atoms/mol x 3.000 moles = 1.807×10^{24} atoms

Calculating moles of molecules

Remember that to calculate relative molecular mass (this is the name we use for molecules), all you do is add up the relative atomic masses of all the different atoms in the molecule. What could be simpler?

When finding the mass of 1 mol of a compound, you follow the same procedure as for finding the mass of 1 mol of an element. Work out the relative molecular mass and convert to grams — this gives you the mass of 1 mol of that compound.

For instance, what is the mass of one mole of methane gas (CH_4)?

1. Write the molecular formula CH_4. Remember the subscripts tell us the number of atoms of each type of atom in the molecule, in this case, C and H.
2. Each molecule is made up of 1 C-atom and 4 H-atoms. The relative molecular mass is [12.01 + 4(l .008)] =16.042
3. The molar mass of CH_4 is thus 16.042 g mol^{-1}

Activity 7.9 will give you some quick practice before you carry on.

Activity 7.9

What is the mass of 1 mole of

a. Calcium carbonate b. Silicon dioxide

Guided reflection

In order to do this activity, you need to make sure you have the correct formula.

You learnt how to do this in unit 5.

Table 1 in unit 5 will help you. You have to know the charge on the calcium ion and the carbonate ion.

If you look up carbonate, you will find it is CO_3^{2-} and calcium is Ca^{2+}. Luckily, they both have equal but opposite charges, so the formula is $CaCO_3$. So calcium carbonate has on Ca atom, one C atom and three O atoms. If you consult the periodic table you will find that one mole would have a mass of 40.08 + 12.01 + (3 x 16.00) = 100.09.

For silicon dioxide, you will have to work it out because it is not on table 1 in unit 5. Two things can guide you. Firstly the letters *di* in front of *oxide* means there are two O –atoms, so the formula is SiO_2. Also, you know Si is in group 4, which means it will share 4 electrons like carbon so it will form SiO_2 in the same way that carbon dioxide is CO_2.

So the mass of one mole of SiO_2 is [28.09 + (2x16.00)] = 60.09 g

By now you probably realise that if you have 1 mol of methane (CH_4) in a container then if you could count the molecules you would find that you have an amount of 1 mol of molecules. However, something that many people find confusing is the idea since each molecule is made up of 1 C- atom and 4H-atoms, you also have 1 mol of C- atoms and 4 moles of H-atoms.

It is very difficult to visualize a mole of anything because the mole is such a big number. It is easier for us to think in dozens. If something is true about a dozen items, then this can be extended to the bigger number, which is the mole. A simple example follows. Think of 1 dozen faces.

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The following things are true about this number of faces. There 1 dozen (12 faces), 1 dozen mouths (12 mouths), 2 dozen eyes (24 eyes).
You can see that even when the number is expressed in dozens, the number of eyes is twice the number of faces! One dozen faces but 2 dozen eyes.
The same is true for 1 mole of faces. it is impossible to draw one mole of faces but here is a drawing which just shows that a mole is a large number.

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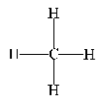
☺ ☺ $\times 10^{22}$

If you have 1 mole of faces (6.022 x 10^{23} faces) then there are 1 mole of mouths (6.022 x 10^{23} mouths) and 2 moles of eyes (2 x 6.02 x 10^{23} eyes).
The same is true for molecules! Here you focus on the subscripts in the formulae to tell you the relationship between moles of molecules and the moles of each different types of atoms.

CH_4 or 

https://www.google.co.za/search?q=image+CH4&tbm=isch&source=hp&sa=X&ved=2ahUKEwj89_GE99PgAhVlt3EKHXofCiUQsAR6BAgDEAE&biw=1536&bih=736&dpr=1.25#imgrc=jwdGGnVv5wJacM:
1 mole of CH_4 molecules (6.022 x 10^{23} molecules) has 1 mole of C -atoms (6.022 x 10^{23} C-atoms) and 4 moles of H-atoms (4 x 6.022 x 10^{23} H-atoms).

It is time to pull everything together and see how you are getting on. They are very easy to do after some practice.

Summary Assessment Unit 7

For both questions 1 and 2 in this revision you will need to consult table 1 of unit 5 and the periodic table.

Task 1

Write balanced equations for the following reactions:

a. a silver nitrate solution reacts with a sodium iodide solution to form silver iodide which forms a precipitate (a solid substance which does not dissolve) and sodium nitrate in solution.
b. a sodium hydroxide solution reacts with a hydrochloric acid solution to form an sodium chloride solution and water.

c. magnesium reacts with a solution of hydrochloric acid to give magnesium chloride solution and hydrogen gas.

Task 2

Complete the following table. If the formula of the substance is not given, work it out before on start calculating.

Substance	Mass	Amount (moles)	Number of molecules	Number of atoms
hydrogen	2.0 g			H-atoms:
helium	40.0 g			
H_2SO_4	24.5 g			H-atoms: S-atoms: O-atoms:
O_2	4.00 g			

Feedback for summary Assessment

Task 1

For each equation, you first need to find out the formulae of the substance sin the reaction.

a Referring back to table 1 in unit 5, you will see that silver ions have a charge of 1 (Ag^+) and nitrate has a formula NO_3^-, with a charge of -1. Therefore, silver nitrate has a formula of $AgNO_3$. Sodium is in group 1 of the periodic table so it forms an ion with a charge of +1. Iodine is in group 17 and so its ion has a charge of -1. Therefore, its formula is NaI. We also need to work out the formulae of the products. Silver iodide will be AgI and sodium nitrate will be $NaNO_3$. Can you work out why?
So the equation is $AgNO_3$(aq) + NaI(aq) → $NaNO_3$(aq) + AgI (s)
We are lucky here as the equation balances with no further work. We show the precipitate as a solid in brackets.

b. Referring back to table 1 in unit 5, you will see that OH^- has a charge of -1. As before, sodium is in group 1 of the periodic table so it forms an ion with a charge of +1. Therefore, the formula for sodium hydroxide is NaOH. Hydrochloric acid is a compound of H^+ (charge +1) and Cl^- (charge -1) so its formula is HCl. We also need to work out the formulae of the products. sodium chloride will be Nacl and you should know the formula for water which is H_2O
So the equation is NaOH (aq) + HCl (aq) → NaCl + H_2O (ℓ)
Again the equation is balanced as it stands

c. In its elemental form, magnesium has no charge. We have the formulae for hydrochloric acid, which is HCl. Magnesium ions have a charge of +2 which you can work out from the periodic table, and we already know chloride as having a charge of -1.
So the *unbalanced* equation is Mg(s) + HCl (aq) → $MgCl_2$ + H_2(g)
To balance the equation, we need a 2 in front of the HCl. This will give us 2 Cl atoms and 2 H atoms on the right hand side. So the balanced equation is
Mg(s) + 2HCl (aq) → $MgCl_2$ + H_2(g)

Task 2

The complete table with working is shown below:

Substance	**Mass**	**Amount (moles)**	**Number of molecules**	**Number of atoms**
hydrogen	2.0 g	2.0g /2.016 g mol^{-1} = 0.99 mol	0.99 x 6.022 x 10^{23} = 6.0 x 10^{23}	H-atoms: 2 x 0.99 x 6.022 x 10^{23} = 6.0 x 10^{23}
helium	40.0 g	40.0g /4.00 g mol^{-1} = 10.00 mol	0 molecules (He is monatomic)	He-atoms: 10.00 x 6.022 x 10^{23}

Substance	Mass	Amount (moles)	Number of molecules	Number of atoms
				= 6.01 x 10^{24}
H_2SO_4	24.5 g	24.5g /98.06 g mol^{-1} = 0.249 mol	0.249 x 6.022 x 10^{23} = 1.50 x 10^{23}	H:2 x 0.249 x 6.022 x 10^{23} = 6.0 x 10^{23} S: 1 x 0.249 x 6.022 x 10^{23} = 3.0 x 10^{23} O:4 x 0.249 x 6.022 x 10^{23} = 5.99 x 10^{23}
O_2	4.00 g	4.0g /32.00 g mol^{-1} = 0.125 mol	0.125 x 6.022 x 10^{23} = 7.53 x 10^{22}	O:2 x 7.53 x 10^{22} = 1.51 x 10^{23}

Working out the mass of 1 mol of H_2SO_4 – 2 x H (2.0g) + 1 x S (32.06) + 4 x O (16.00)

Key learning points from unit 7

Now that you have completed the study of unit 7, you should be able to:

- Define what is a chemical reaction
- Write simple chemical equations
- Describe one mole of a substance the unit of amount containing Avogadro's number of particles;
- Define molar mass;
- Calculate the molar mass from the chemical formulae of chemical compounds, using the periodic table.

Unit 8: Intermolecular Forces

Outcomes

By the end of this unit, you should be able to:

- Define and differentiate between intramolecular and intermolecular forces;
- Explain the following types of intermolecular forces: ion-dipole forces, ion-induced dipole forces, Van der Waals forces (dipole-dipole, dipole-induced dipole and induced dipole-induced dipole/London/dispersion forces), hydrogen bonds;
- Identify the type of intermolecular force between specific types of particles (atoms, ions or molecules);
- Explain the effect of intermolecular forces on boiling point, melting point and rate of evaporation.

Introduction to Unit 8

This unit deals with the forces between the particles in materials. Like the forces which lead to chemical bonding, these forces are electrostatic which means that they are forces between positive and negative charges. In all atoms there are positive nuclei and negative electrons. In unit 4, we learnt about forces within the atom, that is, forces between the nuclei and electrons. These forces hold the atom together. Then in unit 6, we learnt about bonding which is due to forces *between atoms*. In this unit we will learn about forces *between molecules* which are generally weaker than the forces that cause bonding.

We use the term *intramolecular* to mean **inside** the molecule, and *intermolecular* to mean **between** molecules. Intramolecular forces are the bonds between the atoms that make up the molecule.

These are strong electrostatic forces and can be covalent, ionic or metallic bonds that you learnt about in unit 6.

For example, in the iodine molecule there is a strong covalent bond between each of the two atoms. This strong force pulls the atoms close together which is why we draw the two circles overlapping:

However, in an iodine crystal, there are millions of molecules, and these molecules are held together by weaker intermolecular forces, shown as dotted lines in figure 8.1.

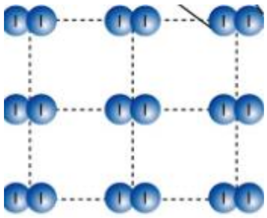


Figure 8.1: Forces holding iodine atoms together (left) to form an iodine crystal (right)
https://slideplayer.com/slide/8729059/
https://www.google.com/search?rlz=1C1GCEA_enZA815ZA815&q=image+iodine+crystal&tbm=isch&source=univ&sa=X&ved=2ahUKEwiBq-SDztTgAhUGTBUIHXjyCHAQsAR6BAgCEAE&biw=1920&bih=969#imgrc=38awHpTyOsjA4M:

How are intermolecular forces related to the world around us?

Have you ever wondered why some substances are solids at room temperature, and other substances are liquids or gases? The answer lies with the intermolecular forces between the particles of the substance. Intermolecular forces also explain why oil floats on water and does not dissolve in it. The most amazing of all is that ice floats on liquid water. Look again at Figure 1.9 in unit 1. For almost every other substance, the solid phase is denser than the liquid phase of the substance, and so the solid phase sinks. So much of our Earth's environment depends on this fact, and the explanation is in terms of intermolecular forces.
Many other properties of materials are based on the intermolecular forces in the material: for example, why alcohol evaporates faster than water, why water moves up the stems of plants as shown in figure 8.2, why insects can walk on water as shown in figure 8.3, or why some materials are brittle and other are strong and hard as shown in figure 8.4.

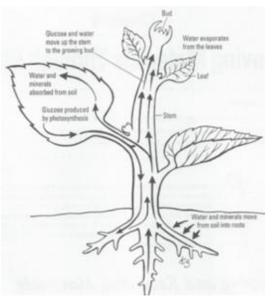


Figure 8.2: Intermolecular forces enable water to move up the stems of plants
http://www.sps186.org/downloads/basic/511123/Introduction,%20Diffusion,%20Osmosis.pdf

Figure 8.3: Intermolecular forces enable insects to walk on water
https://physicsworld.com/a/walking-on-water/

why some materials are brittle and other are strong and hard.

Figure 8.4: Intermolecular forces are responsible for some of the properties of materials
http://www.kefaloniabyanna.com/web/wp-content/uploads/2016/05/pottery-breaking12.jpg

An overview of the forces between molecules

In this unit we will look at the different types of intermolecular forces: dipole-dipole forces, ion-dipole, London forces and hydrogen bonds, and we will identify how these forces arise within both molecular and ionic types of structures. We will show how the shape of the molecule and the distribution of electrons in the molecule affect the way molecules are attracted to each other. We will also look at some of the properties of the materials that are due to the intermolecular forces, such as the melting and boiling points, and rates of evaporation. This is a link to a 10 min video giving an overview of the topic:
https://www.youtube.com/watch?v=08kGgrqaZXA&vl=en

Activity 8.1

First, let's recap what you have learnt in unit 6 about forces *within* molecules. Try not to look back when responding.

1. Choose the correct word:

a. A bond that involves equal sharing of a pair of electrons is called *ionic/covalent*

b. The bonds in a crystal of Cu(s) are *metallic/ionic*

c. Covalent bonding involves the *delocalised/ valence* electrons.

2. Decide whether the following molecules are polar or non-polar and explain your answer in each case: bromine (Br_2), water (H_2O), carbon dioxide (CO_2), ammonia (NH_3)

Guided Reflection

The correct answers are given below:

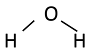
1. a. When a pair of electrons is shared between two atoms it is called a *covalent* bond. This usually happens between non-metals

b. Cu(s) is a metal and so the bonds within the crystal are *metallic.* All the atoms can be from the same element

c. Covalent bonding involves the *valence* electrons

2. Br_2 is non-polar since the two atoms in the molecule are identical.

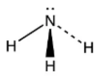
H_2O is a polar molecule. The molecular shape is bent:



Oxygen is more electronegative than hydrogen. This means that the electrons spend more time on the oxygen. Thus the oxygen gives the molecule an electronegative side and the H atoms give the molecule a positive side.

The CO_2 molecule is different. Even though the bonds are polar because O is more electronegative than the C, the molecule is non-polar. This is because the molecule has a linear shape O=C=O, and so the electronegativity of each side of the molecule is the same. The molecule is non-polar. See Unit 6.

NH_3 is a pyramidal shaped molecule,



https://www.quora.com/What-is-the-shape-of-an-NH3-molecule

The N atom is more electronegative than the H atoms so the molecule is polar.

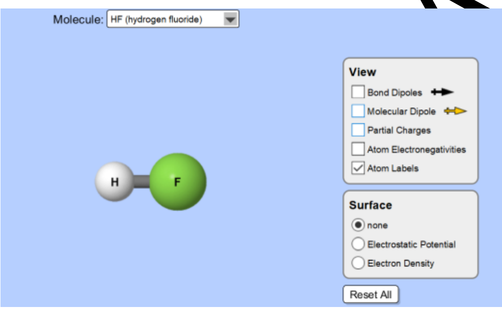
The following activity will help you see this more clearly.

Activity 8.2

Look at this simulation to see the polarity of molecules.

https://phet.colorado.edu/en/simulation/legacy/molecule-polarity

Open up the simulation and select Real Molecules. You will see the view below:



Under "view" select all the options (bond dipoles, molecular dipole, partial charges, atom, electronegativities and atom labels). Under "surface" select "none". Now try the following molecules.

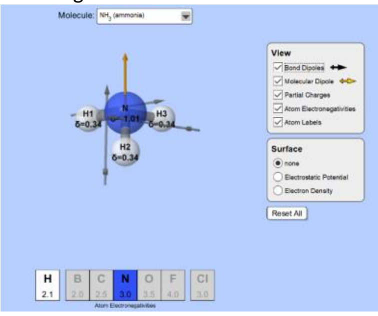
1. Try two of the molecules where both elements are the same (two of H_2, N_2, O_2 F_2). What do you notice about the difference between the electronegativities? Try clicking and unclicking the other options under "view". Does anything change?

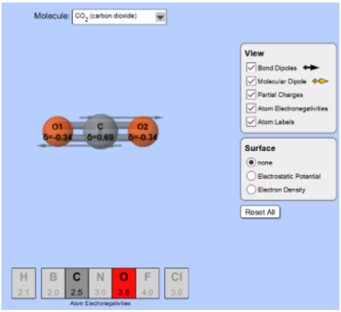
2. Now try the other molecules mentioned in activity 8.1 number 2. They were water (H_2O), carbon dioxide (CO_2) and ammonia (NH_3). Unclick all the options under "view" except "atom labels" and atom electronegativities" and then click one at a time. What do you see? What differences are there for the different molecules?

Guided Reflection

1. All the molecules where both elements were the same gave similar answers. All showed partial charges as δ = 0, meaning that the difference in electronegativity was zero. All showed no bond or molecular dipoles.

2. With water (H_2O), carbon dioxide (CO_2) and ammonia (NH_3), you will notice that water and ammonia were similar. In both cases the partial charges showed positive values for hydrogen and negative values for the nitrogen or oxygen. The bond dipole showed polar bonds in all cases and the molecular dipole showed that the molecule was polar. But carbon dioxide was different. Partial charges were positive for the carbon atom and negative for the oxygen atoms. But the one big difference was that there was no bond dipole, confirming what we saw in activity 8.1. This shows again that not all molecules with polar bonds are polar. It depends on the shape of the molecule. the following two views show the difference:





A reminder about electrostatic forces

You will remember from your physics module that like charges repel and unlike forces attract. This means, for example, that two positive charges will repel each other, and a positive and negative charge will attract one another.
The strength of the force depends on the strength of the charge and the distance between the charges. As the charges get closer together, the forces increase. You may remember this formula from your physics units:

$$F_e = k_e \frac{q_1 q_2}{r^2}$$

where

- F_e is the force
- k_e is the Coulomb's constant (8.987×10^9 $N.m^2.C^{-2}$)
- q_1 and q_2 are the signed magnitudes of the charges
- r is the distance between the charges

https://www.quora.com/What-happens-to-the-electric-force-between-two-objects-when-the-distance-between-them-increases-Decreases

Figure 8.2: Coulomb's law

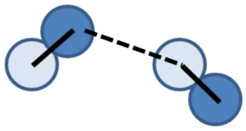
In this application to intermolecular forces, it means that the sizes and shapes of the molecules are important. Symmetrical molecules like carbon dioxide are the same on both sides. Non-symmetrical molecules like ammonia are not the same on both sides, and so the distribution of electrons differs and the molecule is polar. There is an attractive force between the charge on the more positive end of one molecule and the more negative end of another molecule. This force holds the molecules together. So molecules with polar bonds may or may not be polar depending on the symmetry of the molecule.

Two important effects of Coulomb's Law are important to chemistry:
The kinetic theory that we learnt in unit 1 assumes that there are **no** forces between the molecules of a gas. The gas molecules are far apart and thus the force between them becomes very small and can usually be ignored. However, in liquids and solids the particles are much closer together, and the forces between the particles are stronger. When we discuss intermolecular forces we will deal only with the solid and liquid phases.

Another aspect to consider when we apply Coulomb's Law, is that *within* a molecule, the charges are close together and so the bonds are strong. The distances *between* molecules are larger and thus the intermolecular forces are much weaker than bonds.

Activity 8.3

Figure 8.3 shows a model of two molecules in a liquid state:



What do the solid lines represent?
What does the dotted line represent?
Compare the strength of these and explain why they are different.

Guided Reflection

The solid lines represent bonds between atoms inside a molecule.
The dotted line represents an intermolecular force between the molecules.
The bond is a stronger force because the charges are closer together. (The dotted line is longer.) The intermolecular forces are weaker because the distance between the molecules is greater than the distance between atoms inside the molecule.

Activity 8.4

You will need a 500 ml glass bottle, a glass marker pen, a funnel, a glass 250 ml measuring cup or cylinder, 750 ml water and 250 ml of methylated spirits. *Note that methylated spirit is poisonous so do not inhale it.*
Carefully measure out 250 ml of water and use the funnel to pour it into the bottle. Add another 250 ml of water. Mark the level of the water. Now empty the bottle and discard the water.
Measure another 250 ml of water and pour it into the bottle. Measure 250 ml of methylated spirits and add it to the water in the bottle.
Shake and observe the level of the liquid. Can you explain the difference in levels?

You can watch this experiment here: https://www.youtube.com/watch?v=b06lm7MEmY8
alternatively https://www.youtube.com/watch?v=L1-Ay5JYi10 (Same video)

Guided Reflection

We find that the level of the mixture of water and methylated spirits is less than the level of pure water. This means that the volume of the mixture is less than the volume of the two separate liquids.
The reason for this is that there are spaces between the water molecules, and there are attractive forces between the water and methylated spirit molecules. When the alcohol molecules are attracted to the water molecules, they move closer to each other due to the attraction.
Now that we have learnt something about inter and intramolecular forces, let us look at the different types.

Types of intermolecular forces

There are basically the following types of intermolecular forces: van der Waals forces, hydrogen bonding, and ion-dipole forces. First we will discuss van der Waals forces which are divided into dipole-dipole, dipole-induced dipole and London forces.

Dipole-dipole forces

Above we looked at polar molecules like ammonia. Because these molecules are polar, they can be described as dipoles.
A dipole is a separation of the positive and negative charges, similar to the separation of magnetic poles in a bar magnet. Polar molecules have a permanent dipole, this means that the differences in electronegativity is always there. Polar molecules could consist of two unlike atoms, for example, HCl or CO, or the molecule could have a shape that causes it to be non-symmetrical, e.g. SO_2. A substance consisting of polar molecules has attractive dipole-dipole forces between molecules. Since the one end of the molecule is slightly positive (shown in figure 8.4 as δ+) and the other is slightly negative (shown in figure 8.4 as δ-), The δ+ end of the one molecule is attracted to the δ- end of the other molecule.

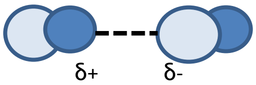
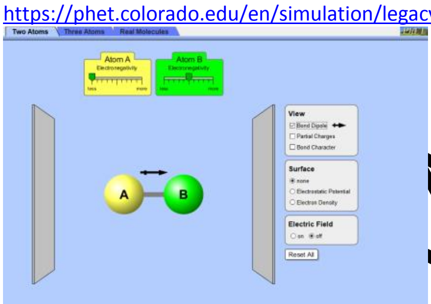


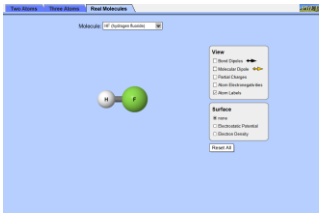
Figure 8.5: Attraction between two dipoles
In activity 8.5 we return to the simulation and use one of the other features which show the electron clouds.

Activity 8.5

Look again at this simulation to see the polarity of molecules:
https://phet.colorado.edu/en/simulation/legacy/molecule-polarity

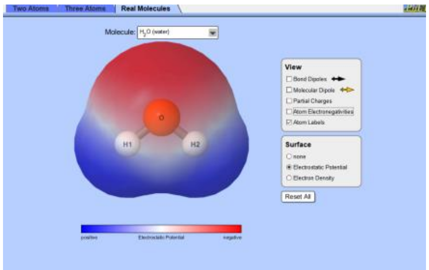


Now click on the **Real Molecules** tab at the top and you will get this screen:



Now try clicking H_2O and NH_3. (Make sure that the box Atom labels is ticked.)

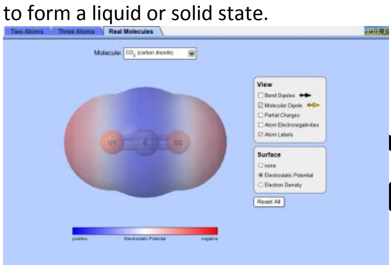
If you tick the box Electrostatic potential, you get diagrams like this one for a water molecule. It shows the positive side of the molecule in blue and the negative side in red.



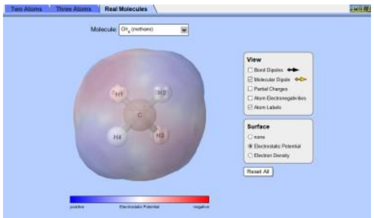
1. Now click through other molecules in the list and find three non-polar molecules. Explain how you know they are non-polar.
2. Now compare the CO_2 molecule with the CH_4 molecule and use polarity to explain why the melting point of CO_2 is much higher than the melting point of CH_4.

Guided Reflection

1. Non-polar molecules shown are H_2, N_2, O_2, F_2, CO_2, BH_3, CH_4, and CF_4. The sides of the molecules are the same.
2. Although the CO_2 molecule is non-polar, the middle atom is more positive than the other two atoms in this linear arrangement. So the O atom from another molecule could be attracted to the C atom by an intermolecular force. This means that the molecules of the substance can come together to form a liquid or solid state.



However, in the CH_4 molecule there is no negative or positive atom on the outside which could attract another atom



Dipole-dipole forces also exist between mixtures of substances where the molecules are polar. For example, between molecules of ethanol and hydrochloric acid, the positive polar end of the H atom in ethanol is attracted to the negative Cl atom of the HCl as shown in figure 8.5. This is also what happened in the case of the methylated spirits and water in activity 8.4. Methylated spirits is mostly ethanol.

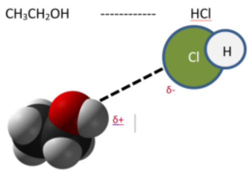
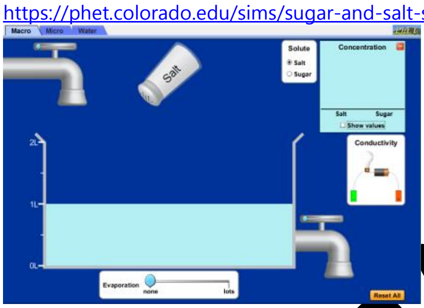


Figure 8.5: Attraction between dipoles HCl and ethanol
https://kids.kiddle.co/images/thumb/0/00/Ethanol-3D-vdW.png/563px-Ethanol-3D-vdW.png
(for ethanol molecule)

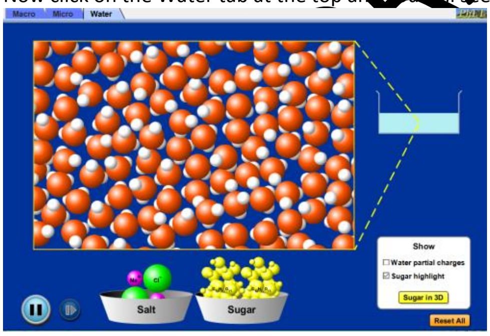
In a polar mixture, the intermolecular forces cause the molecules to align themselves so that oppositely charged particles are attracted.

Activity 8.6

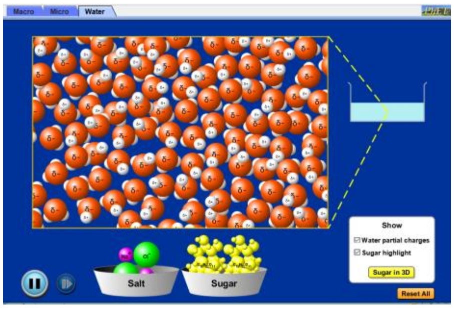
Look at this simulation of how the molecules align themselves in a solution:
https://phet.colorado.edu/sims/sugar-and-salt-solutions/sugar-and-salt-solutions_en.jar



Now click on the Water tab at the top and you will see this model of water molecules:



Now click the box labelled: 'Water partial charges' on bottom right:

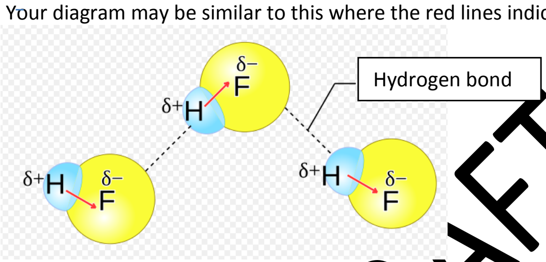


Notice how the opposite charged atoms are attracted to each other.

Now draw a diagram of a few molecules of HF. Label the bonds and the hydrogen bonds.

Guided Reflection

Your diagram may be similar to this where the red lines indicate polar bonds:



https://commons.wikimedia.org/wiki/File:HF_Hydrogen-bonding_illust_jp.svg

Figure 8.6: Forces within and between molecules of hydrogen fluoride

Dipole-induced dipole forces

In a mixture of a polar substance and a non-polar substance, the polar molecule can induce (create) a temporary dipole in the non-polar molecule. As the polar molecule comes close to the non-polar molecule, the negative polar side repels the electrons of the non-polar molecule. The electron cloud distorts to make an induced dipole in the non-polar molecule.

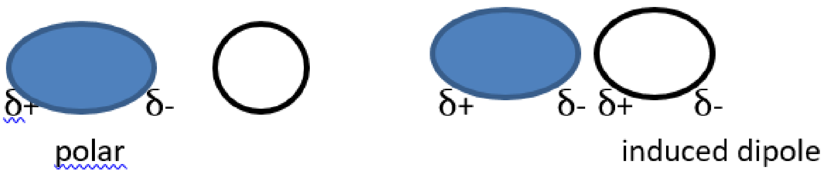


Figure 8.7 Induced polarity in a molecule

For example, oxygen dissolves in water because the dipole in the water molecules induces a dipole in the oxygen molecules. This cause an attraction between the two types of molecules and they disperse between each other.

London forces

London forces are also called induced dipole - induced dipole forces or dispersion forces.

These occur between non-polar molecules. The forces are small and are temporary so we only need to consider them when there are no other forces between molecules.

The electron cloud around a molecule is continually in motion and can momentarily distort. When this happens it can induce another temporary dipole in a neighbouring molecule. Then a temporary attraction occurs between the molecules.

Large molecules or atoms have electrons further from the nuclei and so it is easier to distort the electron cloud than with smaller molecules. For example, iodine atoms are relatively large. The iodine molecule, I_2, is non-polar, but iodine can exist in a liquid or solid state which indicates that there are intermolecular forces between the molecules.

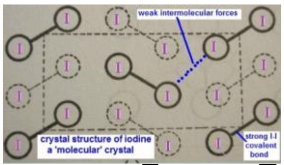


Figure 8.8: Model of an iodine crystal
http://www.docbrown.info/page04/4_72bond64.htm

Dry ice, the solid form of CO_2, is another example of a non-polar substance that forms temporary dipoles. In this model you can see how the C atoms (black) are attached to the O atoms (red).

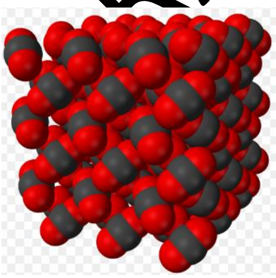


Figure 8.9: Model of molecules in dry ice, $CO_2(s)$
https://commons.wikimedia.org/wiki/File:Carbon-dioxide-crystal-3D-vdW.png

Tetrachloromethane, CCl_4, is a non-polar molecule since the Cl atoms are symmetrically distributed around the central C atom. Iodine dissolves in CCl_4, so this is an example of two non-polar liquids that induce temporary dipoles in their molecules.
We have discussed van der Waals' forces. The next category of intermolecular forces is a very important category known as hydrogen bonding.

Hydrogen bonds

This is a particularly strong *inter*molecular force, even though it is called a bond. It occurs only between molecules that contain hydrogen atoms and any of these three types of atom: N, O, or F.

All of these elements are more electronegative than hydrogen. Water is the most common example. As you know, the water molecule has a bent shape and there is a high electronegativity difference between oxygen and hydrogen, so the water molecule is polar.

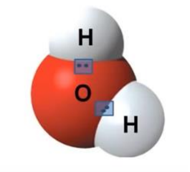


Figure 8.10 A single molecule of water

https://www.youtube.com/watch?v=wZ9CCj4X28Q

A hydrogen bond then occurs between the hydrogen in one atom and the oxygen in another water molecule. We will look at this in more detail later.

We have seen that there is a high electronegativity difference between H and N, O or F atoms. See table 8.1

Table 8.1 Electronegativity of certain atoms

H	N	O	F
2,1	3,0	3,5	4,0

This gives rise to the forces between molecules. Notice that we are not discussing the bond *inside* the molecule, but the intermolecular force *between* the molecules. This is a strong attractive force between a H-atom in one molecule and an N, O or F atom in another molecule and is often called a hydrogen bridge. See diagram 8.11 which shows forces between water molecules.

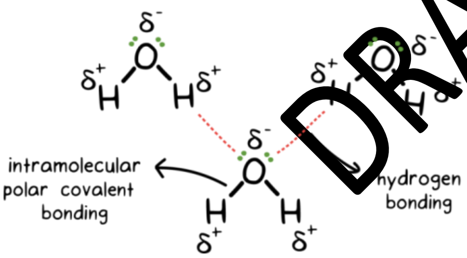


Figure 8.11 Intermolecular forces between H_2O molecules

https://www.khanacademy.org/test-prep/mcat/chemical-processes/covalent-bonds/a/intramolecular-and-intermolecular-forces

The dotted lines indicate the intermolecular forces.

Examples of substances where hydrogen bonding occurs are H_2O, HF, NH_3 and in many organic molecules, such as ethanoic acid, some polymers and even in the DNA molecule as shown in figure 8.12.

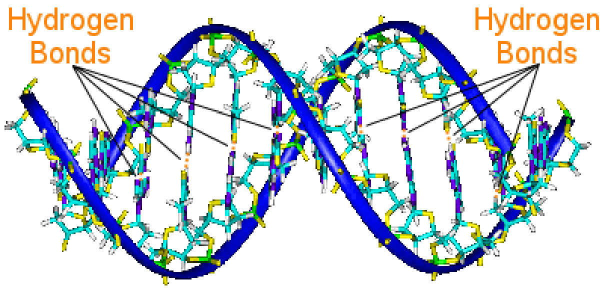


Figure 8.12: DNA molecule showing H-bonds between the chains.
https://www.chemicool.com/examples/hydrogen-bonding.html

This strong force of attraction in hydrogen bonding has important effects on the physical properties of the substances. When a substance melts or boils, the molecules need to separate. If hydrogen bonding is present, much more energy is needed to separate them, and so the melting and boiling temperatures are much higher than expected from their molecular mass. Have a look at figure 8:13 below:

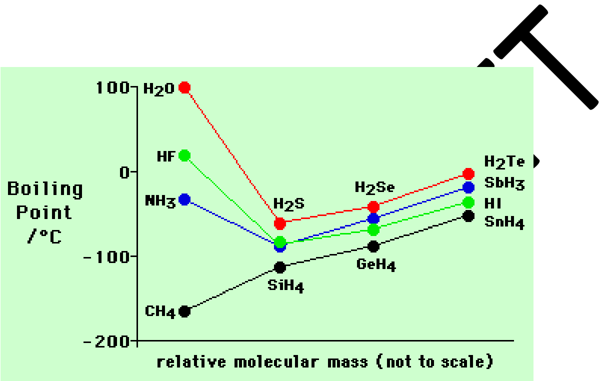


Figure 8.13: Graph of boiling points of similar hydrogen compounds
https://legacy.chemgym.net/environmental_chemistry/topic_3c/page_4.html
The rate of evaporation is also determined by the energy needed to separate the molecules in a liquid, that is, to vaporize the substance. Consequently, compounds containing H-bonds have a much lower rate of evaporation, as you can see in this table below.
Table 8.2 Energy needed to vaporize a substance

Type of substance	**Compound**	**Energy (kJ/mol)**
Hydrogen bonding	HF	25
	NH_3	23
	H_2O	41
	CH_3OH	35
No hydrogen bonding but polar compounds	HCl	16
	HI	20
Non-polar compounds	CH_4	8
	N_2	6
	H_2	1

Activity 8.7

You will need a ceramic tile or flat piece of glass, a dropper, a timer and a few drops of water, acetone, ethanol and methylated spirits. (Safety note: Do not inhale any of these substances)

Place a drop of methylated spirits on the tile and immediately time how long it takes to evaporate (disappear). Do the same with each of the other liquids. Record your results in a table and suggest the order of strength of inter-molecular forces.

Guided Reflection

All of these substances are polar and contain some hydrogen bonding. You should have found the order from shortest to longest time to evaporate to be: acetone, methanol, ethanol, water. This is an indication that acetone has the weakest intermolecular forces and water has the strongest.

Effects of hydrogen bonding in water

Our environment depends on unique properties of water which are a direct result of the hydrogen bonding.

In the solid state, ice, the water molecules form a regular open-cage structure where the H-atom of one molecule is attracted to the O-atom of another molecule. Four molecules of H_2O surround each H_2O in a tetrahedral structure.

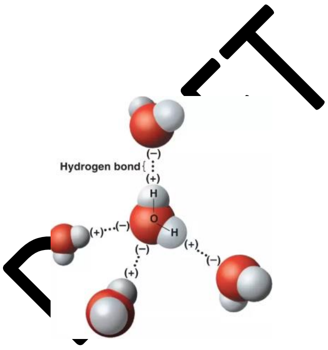


Figure 8.14: Arrangement of molecules in ice

https://www.quora.com/What-is-meant-by-hydrogen-bonding-What-are-the-factors-that-influence-the-formation-and-strength-of-intermolecular-hydrogen-bonding

This open structure in ice is the reason why ice has a lower density than liquid water. The consequence is that ice floats on water and in lakes and oceans, fish and other living organisms survive underneath.

As the temperature rises, the formal structure of ice breaks up and the amount of hydrogen bonding decreases.

This video represents what happens at the molecular level when the ice melts:
http://www.middleschoolchemistry.com/multimedia/chapter2/lesson5

(When accessing this video, you will find that there are several short videos. When you have watched the first one, click on the right arrow to move to the next one. These videos do not have sound)

Because so much energy is needed to break the ice structure, large blocks of melting ice affect the weather and generally moderate temperature of the planet. You will find a similar effect when liquid water evaporates. After you have done some exercise, your body temperature will rise and you will feel hot. Sweating helps to reduce the temperature by evaporation. So you will feel cooler.

You will see the opposite effect when water condenses, so, for example, in a rain storm, energy is released and the temperature rises.

Activity 8.8

Measure the temperature of an ice-water mixture as it is heated.
You will need a thermometer that can measure between 0°C and 25°C, ice, a 250 ml pyrex glass container and heating source, such as a stove or burner.
(Pyrex glass will not crack when you heat it. You can also use an oven proof container. It must not be made of metal. You can buy a thermometer from the swimming pool department at a supermarket.)

Before you start, make a table of time of heating and the temperature as follows:

Time of heating (sec)	**Temp (°C)**
0	
30	
60	
90	
etc…	

Note that the time of heating is related to the amount of energy you put into the system, provided that the heating source remains constant. This means that you need to heat up your heating plate before you start so that it stays the same throughout the experiment. Remember in your unit on the nature of science, you learnt about *variables*. The rate of heating is a *controlled variable.* This means that it must stay the same throughout the experiment. The time of heating is the *independent variable* and the temperature is the *dependent* variable.
Place about 100 ml water in the container and add a few ice cubes. You need to add enough ice cubes to cover the thermometer. Record the temperature in °C. (Some thermometers have two scales, labelled C and F. you need to use the C scale. This stands for Celsius. The Celsius units are written as °C.)
Then heat the mixture and record the temperature every 30 seconds until all the ice melts and continue heating until the temperature is approximately 20°C. You can use a timer on your cell phone, if you have one.
Draw a graph of your results. Use the X axis for your independent variable (time) and the Y axis for the dependent variable (temperature).
Explain why the temperature did not go up at a constant rate. Use the information about hydrogen bonding in your explanation.

Guided Reflection

Your graph should have shown that the temperature remained fairly constant while the ice was melting, even though you were heating the system. Once all the ice has melted, the temperature rises at a constant rate. The temperature did not rise while the ice was melting because the energy from heating was used to separate the ice molecules by breaking some of the hydrogen bonds in the ice structure. Figure 8.15 below shows a typical graph. Notice how the line stays close to zero for quite a long time.

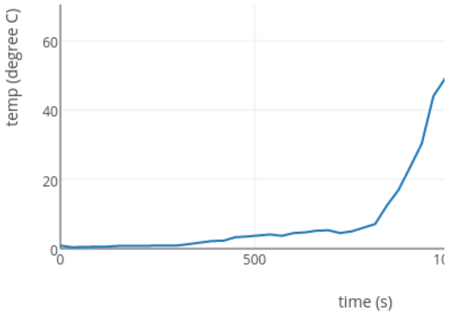


Figure 8.15: Graph of ice melting
https://www.google.co.za/imgres?imgurl=https%3A%2F%2Fplot.ly%2F~17bell_n%2F86%2Frelationship-between-time-and-temperature-of-ice-melting-to-boiling-point.png&imgrefurl=https%3A%2F%2Fplot.ly%2F~17bell_n%2F86%2Frelationship-between-time-and-temperature-of-ice-melting-to-boiling-point.embed&docid=Q5sfqTVYRkMuKM&tbnid=p6C-tHDQp03PiM%3A&vet=10ahUKEwiil4r6pvfgAhVGxYUKHV1CCAwQMwhHKAkwCQ..i&w=700&h=500&bih=881&biw=1745&q=graph%20of%20melting%20ice&ved=0ahUKEwiil4r6pvfgAhVGxYUKHV1CCAwQMwhHKAkwCQ&iact=mrc&uact=8

Ion – dipole forces

You are familiar with the fact that table salt dissolves in water. Table salt, NaCl(s), is made up of Na^+ ions and Cl^- ions packed in an orderly crystal with ionic bonding between the ions. In order to dissolve, the ions need to separate. Because the water molecules are polar, the attractive forces of the different sides of the water molecules are enough to separate the ionic crystal. The water molecules move in between the ions, and ion - dipole forces exist between the water and the salt ions. Notice how the water molecules face in different directions in these two diagrams. We say the molecules are *oriented* in different directions:

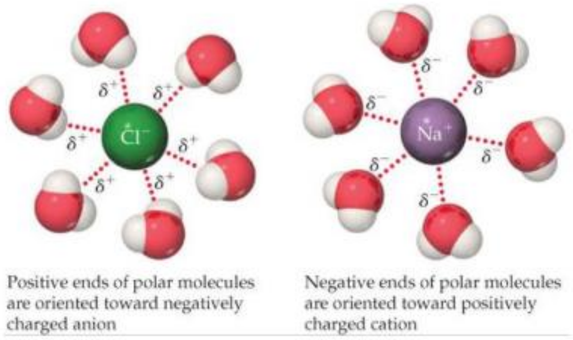


Figure 8.16 Arrangement of molecules and ions in table salt solution
http://forestchemistry.weebly.com/unit-15---intermolecular-forces--bonding.html

Ion – induced dipole forces

When an ion approaches a non-polar molecule, the electron cloud around the non-polar molecule distorts. The non-polar molecule develops partial charges on opposite sides, that is, it has an *induced dipole*. This dipole is temporary and will go away if the ion is removed.

This video explains the difference between ion-dipole and ion-induced dipole interaction and shows the effect of a Zn^{2+} ion in close contact with a non-polar hydrogen molecule (H_2).
https://www.youtube.com/watch?v=DwNmEc9x1U4

Interatomic forces between atoms

We have seen that there are intermolecular forces between molecules and between ions and molecules. However, it is possible for there to be interatomic forces between single atoms such as Helium. Helium does exist in the liquid state at very low temperatures which indicates that there are very weak forces between the helium atoms.

Summary Assessment Unit 8

Task 1

List 4 of the intermolecular forces discussed in this unit. Give an example of a substance that has each of the intermolecular forces you listed.

Task 2

(a) Explain why the boiling point of a substance is dependent on the strength of the intermolecular forces between the molecules.
(b) Put these substances in order of increasing boiling point and explain why you chose that order: HF, Cl_2, CH_3Cl.

Task 3

The specific heat capacity of a substance is the amount of heat energy needed to raise the temperature of 1 gram of the substance by $1^{\circ}C$.
Here are some values for specific heat capacity.

Substance	Specific heat capacity (J)
Water H_2O	4
Ethanol CH_3CH_2OH	3
Tetrachloromethane CCl_4	0,9

Use intermolecular forces to explain the difference in these values.

Task 4

Oxygen dissolves in water – a necessary process for sustaining aquatic life. Explain how a non-polar substance like oxygen can dissolve in a polar substance like water.

Feedback for Summary Assessment

Task 1

The four forces are:
Dipole-dipole e.g. SO_2 sulfur dioxide
Ion-dipole e.g. solution of table salt in water
London e.g. N_2 nitrogen gas
Hydrogen e.g. NH_3 ammonia

Task 2

(a) When a substance boils, the molecules of the liquid separate from each other and leave the liquid and move away form the liquid into the gaseous phase. The molecules in the gaseous phase are much further apart than in the liquid. You remember this form unit 1. If the forces between the molecules in the liquid (intermolecular forces) are strong, more energy is needed to separate them and therefore the temperature at which it boils will be higher.

(b) Cl_2 has the lowest boiling point because the molecules are non-polar and there are weak London forces between the molecules. So when liquid chlorine is heated, less energy is required to separate the molecules.
CH_3Cl has polar molecules and therefore has dipole-dipole forces between the molecules. These forces are stronger than the London forces.
HF has hydrogen bonding between the molecules. These are stronger than dipole-dipole forces and so this substance has the highest boiling point.

Task 3

Water has strong hydrogen bonds between the molecules. When the water is heated, its molecules gain kinetic energy and they move faster and slightly further away from each other (water expands when heated). If the intermolecular forces are strong as in the case of water, it is more difficult to move the molecules apart; more energy is needed and therefore the specific heat capacity is greater.
Ethanol also has some hydrogen bonding due to the O atom in the O-H bond which is attracted to the H atom in a neighbouring molecule. The hydrogen bonding is not quite as strong as in water and the boiling point of ethanol is less than the boiling point of water. The specific heat capacity will be less than that of water because not as much energy is needed to separate the molecule as the substance expands.
Tetrachloromethane is a non-polar substance and has weaker London forces between molecules. Therefore, it will be easier to separate the molecules and require less energy. The specific heat capacity is therefore lower

Task 4

The water molecule has a strong dipole due to its bent shape. When an oxygen molecule is close to a water molecule, the electron cloud around the oxygen molecule becomes distorted and a temporary dipole forms. The side of the oxygen molecule nearest to the hydrogen of the water molecule becomes more negative. The side of the oxygen molecule nearest to the oxygen of the water molecule becomes more positive. This is called an induced dipole and the resultant forces of attraction are called London forces. The oxygen molecules are attracted and can therefore position themselves between water molecules, i.e. the oxygen dissolves in the water. Because the molecules are in motion all the time, the induced dipole keeps changing and new dipoles keep forming.

Key learning points from unit 8

Now that you have completed the study of unit 8, you should be able to:

- Define and differentiate between intramolecular and intermolecular forces;
- Explain the following types of intermolecular forces: ion-dipole forces, ion-induced dipole forces, Van der Waals forces (dipole-dipole, dipole-induced dipole and induced dipole-induced dipole/London/dispersion forces), hydrogen bonds;
- Identify the type of intermolecular force between specific types of particles (atoms, ions or molecules);
- Explain the effect of intermolecular forces on boiling point, melting point and rate of evaporation.

Unit 9: Organic molecules

Outcomes

By the end of this unit, you should be able to

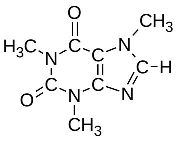
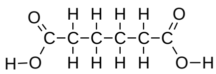
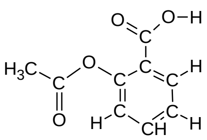
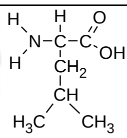
- define the term 'organic' as used in chemistry.
- define the terms functional group, hydrocarbon and homologous series.
- draw the structural, semi-structural, condensed and molecular formulae of organic molecules containing the following functional groups: alkanes, alkenes, alkyne, alkyl halides, alcohols, ketones, aldehydes.
- provide the IUPAC names for organic molecules with up to a maximum of 8 carbon atoms on the parent chain, and a maximum of three functional groups, only alkyl groups or halides.
- identify everyday examples and uses of these various types of organic molecules.

Introduction to unit 9

In the previous units we focused on the atom and the arrangement of atoms on the Periodic Table. We also looked at interactions between atoms (chemical bonds) and between molecules (intermolecular forces). In this unit we are going to look at the molecules that make up almost everything around us – plants, animals, the food we eat, the clothes we wear, cosmetics, medicines and plastics. All these things are made of carbon atoms and form part of a large group of compounds called organic molecules. Here we will only look at some of the basic molecules, how we name them, and why they are important to us.

Why are organic compounds important?

Organic chemistry is the study of organic molecules. As mentioned above, all organic compounds have one thing in common. They all have carbon in them. There are millions of naturally occurring organic molecules, as well as millions of man-made (synthetic) organic molecules. All these play an important role in our daily lives. Organic molecules make up most of our bodies, the clothes we wear, the medicines we take and the food we eat. Figure 9.1 shows some examples of organic molecules in everyday life with their chemical formulae.

  https://upload.wikimedia.org/wikipedia/commons/c/c5/Roasted_coffee_beans.jpg (ChemSketch structure own drawing) *a) caffeine is found in coffee*	  https://upload.wikimedia.org/wikipedia/commons/c/c1/Cardigan%2C_sweater.JPG (ChemSketch structure own drawing) *b) polyester is used to make clothes*
  https://commons.wikimedia.org/wiki/File:201707_medicine_tablets_oval.svg	  https://upload.wikimedia.org/wikipedia/commons/e/ee/Cut_up_steak.jpg

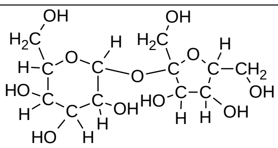
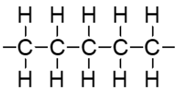
c) aspirin is used in headache tablets	https://commons.wikimedia.org/wiki/File:Leucine.png *d) leucine is found in meat*
  https://upload.wikimedia.org/wikipedia/commons/a/a4/White_cane_sugar.JPG e) *sucrose is sugar*	  https://commons.wikimedia.org/wiki/File:Plastic_Water_Bottle.jpg f) *polyethylene is used to make plastic bottles*

Figure 9.1 Organic molecules in everyday life

It is important to know what these compounds are, and to understand how they are formed, as this can, for example, help us develop new materials for clothing or better medicines to help people.

Activity 9.1: Comparing structures

Have a look at chemical compounds drawn in Figure 9.1. Look at their chemical formulae. What do you notice about their structures? As we mentioned above, they all contain carbon. Count how many carbon atoms there are in each structure. What other elements can you identify in the structures?

Activity 9.1: Guided reflection

You will have noticed that all the compounds in Figure 9.1 have carbon and hydrogen atoms in them, all except polythene (f) have oxygen, leucine and caffeine molecules also have nitrogen, but there are no other atoms involved. Three of the compounds have ring structures. The structures are also quite large, much larger than the ionic compounds and small molecules we have seen to date.
(a) has 8 C atoms, (b) has 6 C atoms. (c) has 8 C atoms, (d) has 6 C atoms, (e) has 12 C atoms and (f) has 5 C atoms. Polyethylene is most interesting because you will see that it has a repeating pattern and is known as a polymer. In reality there are far more than 6 C atoms. You will see more about this in figure 9.2.
Learn more about organic molecules by clicking on: https://youtu.be/stV0Aiqz61Q
(At this point, pay particular attention to the first two and a half minutes)

Carbon is special

Organic molecules have one important thing in common - they are all made of carbon atoms. As you have seen in Activity 9.1 all the compounds have carbon atoms in their structures, and not just one carbon atom, but lots of them! Only carbon form compounds in this way, none of the other elements can do this. What is it that makes carbon special? Carbon has unique properties which influence how it behaves and how it bonds with other atoms.

Carbon has four valence electrons.

Carbon is element 6 on the Periodic Table and has **four valence electrons**. You have learnt about valence electrons in unit 4. Remember valence electrons are the electrons in the outermost energy level. Carbon has four valence electrons, which means it can bond with four other atoms. Carbon atoms often form bonds with hydrogen, oxygen, nitrogen, sulfur and the halogens. This means that a large variety of compounds can be formed.

Carbon can form straight, branched and cyclic molecules

Carbon can bond to four different atoms, as seen above, but it can also bond to other carbon atoms. This means it if carbon just keeps on bonding to another carbon, it can form infinitely long chains. Figure 9.2a shows an example of a soap molecule. Notice that the chain of carbon atoms forms a long tail. The chains can also be branched chains like in Figure 9.2b or link up with one another to form cross-links, like in Figure 9.2c. Lastly, carbon can form ring structures like in Figure 9.2c. This is the plastic PET that is used to make plastic water bottles.

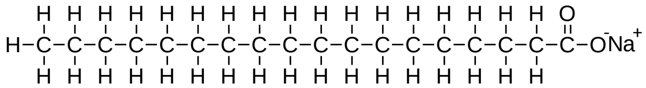
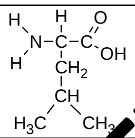
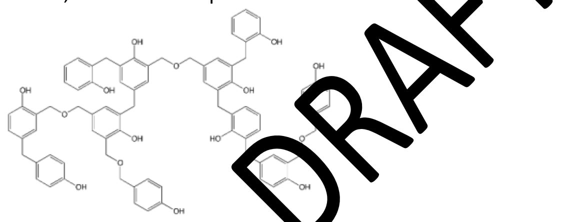
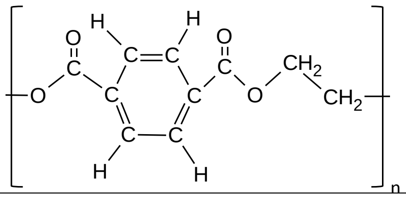
a) sodium stearate in soap 
b) leucine, an amino acid 
c) Bakelite, a resin used in paints  https://commons.wikimedia.org/wiki/File:Phenolic-resin-base-catalysis-5.png
d. d) polyethylene terephthalate (PET) in plastic bottles 

Figure 9.2 Carbon can form long chains, cyclic compounds and branched structures (See above)

Carbon can form single, double and triple bonds

To add to the variety of molecules that carbon can form, it can also form single, double or triple bonds with other carbon atoms, and single and double bonds with oxygen.

Activity 9.2 Identifying double and triple bonds

Identify the single and double bonds that carbon makes in Figure 9.2d by drawing rings around them.

Activity 9.2 Guided reflection

How many double and triple bonds could you identify? Compare your answer to the one I got in Figure 9.3. If only one of these bonds change, you have a new molecule. Can you imagine how many molecules can be formed?

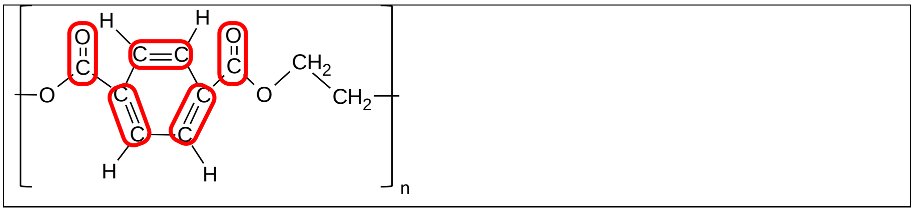


Figure 9.3 Double and triple bonds

Carbon can form polar or non-polar bonds

Most of the bonds that carbon atom form with other atoms are covalent in nature. Carbon is a non-metal and on the right-hand side of the Periodic Table. Think for example of a C – C bond. The difference in electronegativity between the two atoms is zero, and the bond can be classified as a pure covalent bond. In the case of a C – H bond, the difference in electronegativity between carbon (2.5) and hydrogen (2.1) is so small that C – H bonds are almost purely covalent. The result of this is that most organic compounds are non-polar. This affects some of the properties of organic compounds.

However, when carbon bonds to oxygen, the electronegativity difference between the two atoms is more (3.5 for oxygen and 2.5 for carbon). This results in a polar bond. If you look at the soap molecule in Figure 9.2a you will see that the long chain on the left has only C – C and C – H bonds. This 'tail' will be non-polar. The right-hand side, or 'head' of the molecule, has C – O bonds and will be polar. The polarity of the bonds gives molecules unique properties. We will discuss this in the next unit.

Organic compounds are different from inorganic compounds

Where do all the organic molecules come from? In all the chemistry units, so far you have learned about inorganic compounds. Metals react with non-metals to form ionic compounds or non-metals reacting with non-metals to form inorganic compounds. All these compounds are made of small molecules or simple ionic structures. Some of them even contain carbon, for example carbon dioxide (CO_2) or sodium cyanide (NaCN), but they are not organic compounds.

Organic compounds are classified as the compounds of carbon, usually bonded to hydrogen, other carbon atoms, oxygen or nitrogen. It does not include the very simple molecules like carbon dioxide (CO_2) and polyatomic ions like cyanide (CN^-) or carbonate (CO_3^{2-}). These are considered inorganic compounds. You can watch a short video clip on the difference between organic and inorganic compounds here: https://youtu.be/lC57cJzM8OA

The main source of the carbon in organic compounds is carbon dioxide in the atmosphere. Plants use sunlight to convert carbon dioxide and water (inorganic compounds) into sugar (an organic compound) through the process of photosynthesis: $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$
In Life Sciences you learn how plants can make their own organic molecules through photosynthesis, while animals feed on plants or plant products to gain the organic compounds that they need to survive. Organic molecules therefore play a very important role in life on Earth.

Representing organic molecules

There are different ways in which we can represent organic molecules. Organic molecules are three-dimensional structures. When we draw them on a two-dimensional page, it is difficult to capture it accurately. We therefore use different ‘pictures’ or representations of them for different purposes. The one that we will use most often is called the structural formula.

Structural formulae

Language tip: ‘Formula’ is the singular form of the word, with ‘formulae’ the plural.

The **structural formula** shows all the bonds between all the atoms in a molecule. Each line represents two electrons covalently shared between the two atoms. Double bonds are indicated with two lines as can be seen in Figure 9.4.

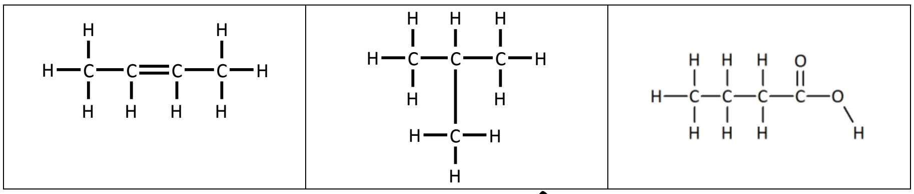


Figure 9.4 The structural formulae for a) but-2-ene b) methylpropane and c) butanoic acid

Semi-structural formulae

You can imagine that is it time-consuming to write out all the bonds, especially if the molecules are large. The semi-condensed structural formula is a shorter way to represent the molecules. Here, we do not write all the C – H bonds, but only the C – C bonds as shown in Figure 9.5. You can compare this figure with the figure above to see how the structures are changed.

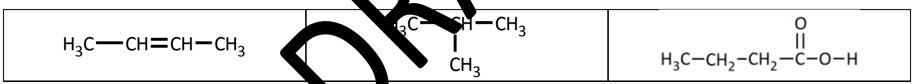


Figure 9.5 The semi-structural formulae for a) but-2-ene b) methylpropane and c) butanoic acid

Condensed structural formulae

We can also represent organic molecules without showing any bonds. Figure 9.6 shows the same three molecules but using condensed structural formulae. This representation is often used in textbooks as it shows enough information but does not take up a lot of space in the text. You will notice that the CH_3 group is in brackets for the second molecule. This is because it is a branched chain which is ‘above’ the main chain. The third molecule, butanoic acid, has an oxygen atom ‘above’ the chain, so we put this in brackets *after* the carbon to which is it bonded.

$CH_3CHCHCH_3$	$CH_3CH(CH_3)CH_3$	$CH_3(CH_2)_2C(O)OH$

Figure 9.6 The condensed structural formulae for a) but-2-ene b) methylpropane and c) butanoic acid

Molecular formulae

Another, even more condensed way of representing organic molecules, is the molecular formula. Here only the type and number of atoms are shown. For example, but-2-ene is C_4H_8 because there are 4 carbon atoms and 8 hydrogen atoms. The molecular formula for methylpropane is C_4H_{10} and for butanoic acid it is $C_4H_8O_2$. Count the atoms in Figure 9.6 to convince yourself that the numbers in Figure 9.7 are correct.

C_4H_8	C_4H_{10}	$C_4H_8O_2$

Figure 9.7 The molecular formulae for a) but-2-ene b) methylpropane and c) butanoic acid

Projection formulae

All the representations so far give the impression that organic molecules are two-dimensional flat structures. This is not the case. Organic molecules are three-dimensional molecules. To show this on a flat page is difficult. The projection formulae help us to create a three-dimensional structure. A line represents a bond flat with the page, a solid wedge represents a bond coming 'out of' the page, and a wedge (looking a bit like a triangle) represents a bond 'into' the page. Figure 9.8 below shows the projection formulae for methane (CH_4) and ethane (CH_3CH_3).

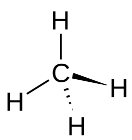
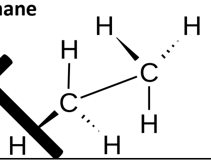
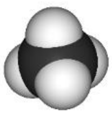
a) methane 	**b) ethane** 

Figure 9.8 Projection formulae for methane and ethane

Three-dimensional molecular models

Organic molecules can also be represented with three-dimensional models. This gives the best special representation. Commercial molecular models are available, or you can build models yourself. Figure 9.9 shows the methane molecule with two kinds of commercial models, ball-and-stick and space-filling molecular models, and a three-dimensional model built with sweets. Try to get hold of some commercial models, look on the internet for pictures of the space filling or ball-and-stick models, or build your own model with sweets like jelly tots and marshmallows, or potatoes and grapes. This will help you 'see' the three-dimensional nature of organic molecules.

]

a) commercial models	b)space filling model
 	 https://commons.wikimedia.org/wiki/File:Methane-3D-space-filling.svg
c) ball-and-stick model	d) marshmallow and jelly tot model

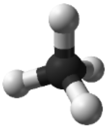
 https://commons.wikimedia.org/wiki/File:Methane-CRC-MW-3D-balls.png	

Figure 9.9: Different Three-dimensional molecular models for methane

Activity 9.3 Building molecules

In this activity you are going to build three-dimensional molecules. You will need marshmallows (representing carbon atoms), toothpicks (representing bonds) and jelly tots (representing hydrogen atoms). Follow the steps below and compare your structures to the ones in the guided reflection below this activity.

1. Use one marshmallow, four toothpicks and four jelly tots to create the molecule in Figure 9.9d. You can cut the marshmallows and toothpick in half for easier handling. Remember carbon must always bond to four atoms, so you always need four toothpicks around carbon. The bonds will also be as far away from each other as in the figure.
2. Write the molecular formula for your structure.
3. Now remove one jelly tot (a hydrogen atom) and replace it with another marshmallow (a carbon atom). Try to keep the angles the same. Add three more toothpicks to the new carbon atom and add three jelly tots to each 'bond'.
4. Write the structural formula for your structure.
5. Now remove another hydrogen atom and replace it with a third carbon (marshmallow). Remember to add three more toothpicks and three hydrogens. How many carbon atoms do you have in your structure?
6. Write down the condensed structural formula for your structure.
7. What is the shape of your molecule?
8. Keep your molecule for later in the unit.

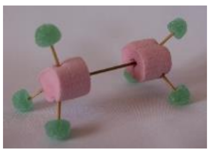
Activity 9.3 Guided reflection

1. I hope you had fun in building these structures. You will notice that the molecule becomes quite large and takes up a lot of space. Compare your structures with the ones below.

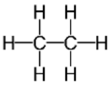
2. CH_4

Remember the molecular formula does not show any bonds, only the different atoms and how many of each. This molecule is called methane.

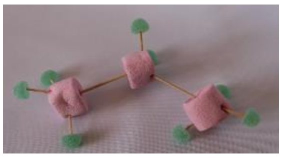
3.



4. Remember that the structural formula shows all the bonds between all the atoms. This molecule is called ethane.



5. There are three carbons in the structure. Does yours look like this?



6. $CH_3CH_2CH_3$ This molecule is called propane.
7. If you managed to keep the angles fixed, you will notice that the molecule is not straight, but at an angle

Functional groups

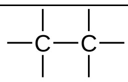
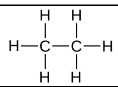
We said earlier in the unit that carbon can form polar or non-polar bonds depending on which atoms it is bonding to. The type and combination of atoms which bond to carbon determine the characteristics of the molecule. This specific combination of atoms is called the functional group.

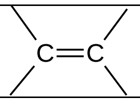
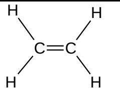
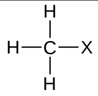
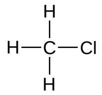
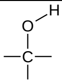
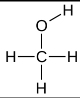
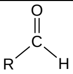
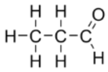
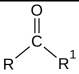
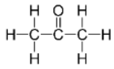
DEFINITION: Functional group
In organic chemistry a functional group is a specific group of atoms (and the bonds between them) that are responsible for the characteristic chemical reactions of those molecules.

For example, in Activity 9.3 all the molecules you built contained only carbon and hydrogen. They are called **hydrocarbons** and are all non-polar in nature and good fuels because they react easily with oxygen. Earlier in the unit, in Figure 9.6c, we had another molecule, butanoic acid, which contains the -C(O)OH group. This is called a **carboxylic acid** group and gives the molecule acidic properties. All molecules containing this group of atoms are carboxylic acids and have acidic properties.

We are now going to look at different functional groups and learn the names of the compounds in each of the groups.

Table 9.1 Different functional groups

Name of the group	Functional group	Name of the functional group	Example	Structural formula
Alkane		Alkane	ethane	

Name of the group	Functional group	Name of the functional group	Example	Structural formula
Alkene		Alkene	ethene	
Alkyne	—C≡C—	Alkyne	ethyne	H—C≡C—H
Haloalkane or alkyl halide	 X=F, Cl, Br, I	Halogen	chloromethane or methyl chloride	
Alcohol		Hydroxyl	methanol	
Aldehyde		Formyl	propanal	
Ketone		Carbonyl	propanone	

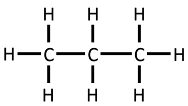
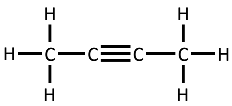
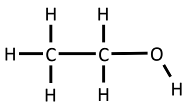
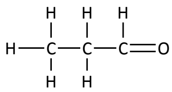
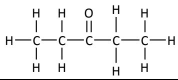
You will notice that the name of a compound contains a first part, called the prefix, and an end part, called the suffix. The prefix indicates the number of carbon atoms according to the following list:

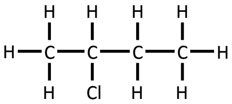
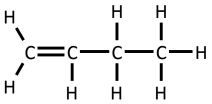
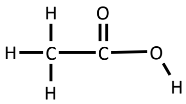
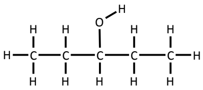
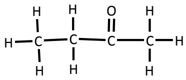
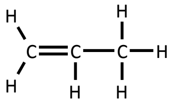
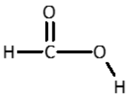
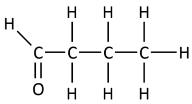
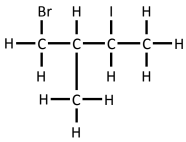
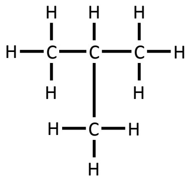
One carbon atom:	*meth-* for example methanol
Two carbon atoms:	*eth-* for example ethene or ethane
Three carbon atoms:	*prop-* for example propanal
Four carbon atoms:	*but-* for example butane

The suffix indicates the functional group and the 'family' or group to which the compound belongs, for example alkanes all end in *-ane*, alkenes in *-ene*, alcohols in *-anol*, etc.

Activity 9.4 Spot the functional group

Let's see if you can identify the functional groups in molecules. Use Table 9.1 to guide you. For each of the examples in the table below, identify the functional group and write down the name of the group to which the molecule belongs. A molecule can have more than one functional group, but in this activity, each one has only one functional group. The first one has been done for you.

1. alkane 	2. 	3. 	4. 
5.	6.	7.	8. 

			
9. 	10. 	11. 	12. 
13. 	14. 	15. 	16. 

Activity 9.4 Guided reflection

Compare your answers to the table below. Be careful to distinguish between an alcohol (numbers 3 and 9) where there is no carbon bonded with a double bond to an oxygen, there is only an hydroxyl (-OH) group, and a carboxylic acid where the carbon is bonded to an -OH group as well as a double bond oxygen, like in number 7 and 13. Also, do not get confused when the molecule is drawn 'upside down', this does not change the molecule.

1. alkane	2. alkyne	3. alcohol	4. aldehyde
5. haloalkane	6. alkene	7. carboxylic acid	8. ketone
9. alcohol	10. alkyne	11. ketone	12. alkene
13. carboxylic acid	14. aldehyde	15. haloalkane	16. alkane

Hydrocarbons

A video lesson on the hydrocarbons can be found here: https://youtu.be/gs1Y7lR8WoI

The hydrocarbons are the simplest organic molecules. They contain only carbon and hydrogen. We will be looking at three different hydrocarbons, the **alkanes** which contain only single bonds, the **alkenes**, which contain double bonds and the **alkynes**, which contain triple bonds.

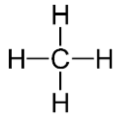
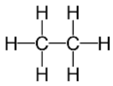
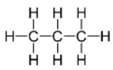
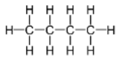
DEFINITION: Hydrocarbon
An organic molecule which contains only carbon and hydrogen atoms with no other functional groups besides single, double or triple carbon-carbon bonds.

Alkanes

The alkanes are hydrocarbons that only contain single covalent bonds between their carbon atoms. The simplest alkane is **methane**. It has only one carbon atom and four hydrogen atoms. The structural, semi-structural and molecular formula is shown below. This the was the first structure you built in Activity 9.3.

The second alkane in the series has two carbon atoms and is called **ethane**. This was the second structure you built in Activity 9.3. The third alkane in the series is called **propane**. It has three carbon atoms and eight hydrogen atoms, all connected with single bonds. The fourth member of the series is **butane** with its four carbon atoms. The table below summarises the first four alkanes in the series. There are many more alkanes as each time a carbon atom is added, a new molecule is formed. As long as the bonds are all single bonds, all these molecules belong to the alkane family.

Table 9.2: The first four alkanes

Name	**Molecular formula**	**Structural formula**	**Semi-structural formula**	**Molecular model** (C is black and H is white and bonds are grey)
methane	CH_4		CH_4	
ethane	C_2H_6		CH_3CH_3	
propane	C_3H_8		$CH_3CH_2CH_3$	
butane	C_4H_{10}		$CH_3CH_2CH_2CH_3$	

Activity 9.5 Examining alkanes

Have a look at the table with the first four alkanes above and compare their structures.

1. Can you see a pattern in the molecular and structural formulae for the alkanes? What is the pattern?
2. Predict what the molecular and structural formulae for the next alkane in the series would be.

Activity 9.5 Guided reflection

Did you find the pattern? Each next alkane has one more carbon and two more hydrogens, so the difference between each alkene and the one before is $-CH_2-$. We can keep on adding a $-CH_2-$ and each time a new alkane is formed. What did you think the next alkane would be? If you have C_5H_{12} and $CH_3CH_2CH_2CH_2CH_3$ you are correct! Can you now see why there are so many organic molecules? Each time one carbon is added, a new molecule is formed.

Uses of alkanes

Alkanes have many important uses. The alkanes are the most important source of fuel in the world and are used extensively in the chemical industry. Alkanes that contain four or less carbon atoms are gases, for example butane. Butane is used in gas cookers. Other alkanes are liquid fuels, for example octane is the main compound in petrol. Long chain alkanes with more than 16 carbon atoms are solids and used in candle wax, shoe polish and many other products.

Gas cooker	Petrol	Wax candle

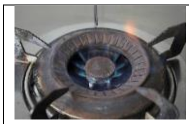
 https://commons.wikimedia.org/wiki/File:Gas_Cooker_(102440303).jpeg	 Find another image or take picture	 https://commons.wikimedia.org/wiki/File:Pillar_candles_09.jpg

Figure 9.10: Alkanes are an important source of fuel

Homologous series

Compounds with the same functional group, which differ from the next one in the series by only a CH_2 group, belong to the same **homologous series**. Compounds in the same homologous series can be represented by a **general formula**. In the case of alkanes, the general formula is $\mathbf{C_nH_{2n+2}}$. The *n* in the formula can be replaced by the number of carbon atoms to get the molecular formula for the specific alkane. For example, an alkane with four carbon atoms will be $C_4H_{2(4)+2}$ to give C_4H_{10}. This is the molecular formula for butane, which is correct since butane contains four carbon atoms and ten hydrogen atoms. The general formula represents all the compounds in a homologous series.

DEFINITION: Homologous series
A homologous series is a series of compounds with the same general formula. All molecules in this series will contain the same functional groups.

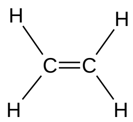
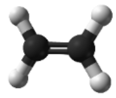
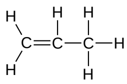
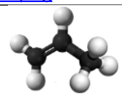
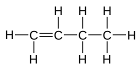
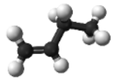
A homologous series is like a family. You belong to a family. All your siblings belong to the same family as you and you may have the same family name (surname). Your family name is like the homologous series, for example alkanes. In your family you and your siblings may have similar features, for example you all have the facial features. These features make you 'belong together'. These features are like the functional groups. All the members of the family have this same feature. All the alkanes have only single bonds and contain only carbon and hydrogen atoms. All alkanes have their names ending in *-ane*.

Alkenes

The next homologous series is the **alkenes**. Alkenes contain only carbon and hydrogen atoms and are therefore hydrocarbons, but alkenes have at least one double bond. The functional group for the alkenes is -C=C- and the general formula is $\mathbf{C_nH_{2n}}$. Table 9.3 below shows the first three alkenes in the series.

Table 9.3: The first three alkenes

Name	Molecular formula	Structural formula	Semi-structural formula	Ball-and-stick model (C is black and H is white)

ethene	C_2H_4		CH_2CH_2	 https://commons.wikimedia.org/wiki/File:Ethane-A-3D-balls.png
propene	C_3H_6		CH_2CHCH_3	 https://upload.wikimedia.org/wikipedia/commons/8/88/Propene3D.png
But-1-ene	C_4H_8		$CH_2CHCH_2CH_3$	 https://commons.wikimedia.org/wiki/File:But-2-ene-3D-balls.png

Uses of the alkenes

The alkenes many uses. Ethene is the compound that causes fruit to ripen, for example green bananas will turn yellow when they come into contact with ethene. Ethene also helps the opening of flowers. Propene is used to make the plastic polypropylene, a strong compound used to make ropes, bottle caps, strong containers and a variety of plastic products.

 bananas https://commons.wikimedia.org/wiki/File:Pl%C3%A1tanos_de_Canarias.JPG	 Polypropylene ropes https://commons.wikimedia.org/wiki/File:Rope_on_ferry.JPG	 Plastic chairs https://commons.wikimedia.org/wiki/Category:Plastic_chairs#/media/File:2011-11-05_wcn_11.JPG

Figure 9.11 Some uses for alkenes

Activity 9.6 Building alkenes

For this activity you will need the marshmallows, toothpicks and jelly tots from Activity 9.3 again.

1. Start with one marshmallow (one carbon atom) and add four toothpicks (‘bonds’). Take care in placing them far apart and in the shape of a tetrahedron.
2. Build the structure for ethene by combining two of the toothpicks and adding one marshmallow. Remember to add hydrogens and make sure that both carbon atoms have four bonds each. Does your structure look like the one in Figure 9.12?
3. Now add another carbon atom (with hydrogens) to make propene. Use the table with alkenes on the previous page to help you. What is the difference between this structure (propene) and the structure for ethene?
4. Now add one more carbon atom (with hydrogens) to make butene. Remember, butene has four atoms. What is the shape of butene, compared to propene and ethene. Do you notice

that you can add the four carbon in more than one place? Draw the structures for each of the four-carbon alkenes that you can make.

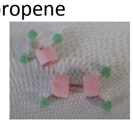
a. ethene	b. propene	c. butene
		 

Figure 9.12 Structures for the first three alkenes. There are two options for a four-carbon alkene.

Activity 9.6 Guided reflection

You will notice that the shape of the structures changes each time another carbon is added. Ethane is a straight molecule, but propene has a kink in it, it is bent. Butene is also bent, either in the middle of the molecule or at the one end.
When you reach a four carbon-atom chain, you would have noticed that there is more than one possible structure. We will look at this later in the unit, but for now you can just make a note that there are more options, each creating a different molecule with different properties. Keep your three structures for later in the unit.

Alkynes

The third group of hydrocarbons is the **alkynes**. Alkynes have at least one **triple bond** between two carbon atoms, and like the alkanes and alkenes, contain only carbon and hydrogen atoms. Look at Figure 9.13a to see how we draw a triple bond. Each line represents two electrons being shared between atoms, so three lines represent a total of six electrons. Do you notice that ethyne has only hydrogen and carbon atoms? Ethyne is the first member of the alkynes and is also known as acetylene. This compound is used in oxy-acetylene blow torches for welding. Alkynes are very reactive, much more so than the alkanes and alkenes. The general formula for the alkynes is $\mathbf{C_nH_{2n-2}}$. For example, propyne has three carbon atoms in the structure, therefore n = 3. The molecular formula for propyne is $C_3H_{2(3)-2} = C_3H_4$.

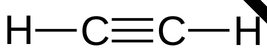
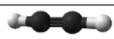
a.	b.
	 b. https://commons.wikimedia.org/wiki/Welding#/media/File:GMAW.welding.af.ncs.jpg

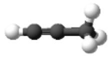
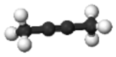
Figure 9.13 Ethyne, also known as acetylene, is a gas that is used for welding

Activity 9.7 Building alkynes

You will need marshmallows, toothpicks and jelly tots, like in Activity 9.6. By now I am hoping that you are comfortable in building structures using these tools. For this activity you need to create a table, just like in Table 9.2 and 9.3. Your table will be for the first three alkynes. Some headings have been created for you below. Build each of the molecules and complete the molecular, structural and semi-structural formulae in the spaces provided in the table.

Table 9.4: The first three alkynes

Name	**Molecular formula**	**Structural formula**	**Semi-structural formula**	**Molecular model** (C is black and H is white)
ethyne				 https://en.wikipedia.org/wiki/File:Acetylene-CRC-IR-3D-balls.png

propyne				 https://upload.wikimedia.org/wikipedia/commons/a/a4/Propyne-3D-balls.png
but-2-yne				 https://en.wikipedia.org/wiki/2-Butyne

Answer the following questions:

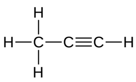
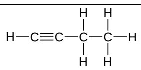
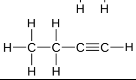
1. Why can you not have methyne?
2. What do you notice about the shape of the carbon triple bond and how this affects the shape of the molecule?
3. What do you notice about the endings of each of the names for alkynes? Where else have you seen this pattern?

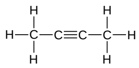
Activity 9.7 Guided reflection

The alkynes follow the same pattern as the alkenes and alkanes, where each next member of the homologous series differs with a CH_2 group from the previous one. Alkanes have single bonds only, alkenes have at least one double bond, and alkynes have at least one triple bond. The names for the alkanes all end in **-ane**, alkenes end in **-ene**, and alkynes end in **-yne**. To answer question 1 above, alkenes and alkynes must have at least two carbon atoms because double and triple bonds cannot form with one carbon atom only. This is the reason why methene and methyne do not exist. To answer the rewound question, Triple bonds cause molecules to be straight, they do not allow the molecule to kink.

Compare the table that you created with the one below. Double check that you have only four bonds on each carbon atom. It is easy to accidentally put more bonds on a carbon atom when you draw the alkynes. Also compare the structures you built with the pictures in the table below. The triple bond causes the molecule to be straight at that point. For butyne there are two options again, we name them but-1-yne and but-2-yne similar to what you had for butene. Later in this unit we will look at why they have different names, and how we decide on the name for a compound.

Table 9.4: The first three alkynes

Name	Molecular formula	Structural formula	Semi-structural formula	Marshmallow and jelly tot model
ethyne	C_2H_4	H—C≡C—H	CH≡CH	
propyne	C_3H_6		$CH_3C≡CH$	
but-1-yne	C_4H_8	 	$CH≡CCH_2CH_3$	

but-2-yne	C_4H_8		$CH_3C{\equiv}CCH_3$	

Alkyl halides

The alkyl halides are hydrocarbons where at least one hydrogen atom is replaced with a halogen, for example fluorine, chlorine, bromine or iodine. The general formula for the alkyl halides is $\mathbf{C_nH_{2n+1}X}$, where X can be any of the halides. Figure 9.14 shows four of the alkyl halides in different representations. Can you see that in each case only one hydrogen has been replaced? In the first one fluorine was used to form methyl fluoride. This molecule is also called fluoromethane. The second one has a chlorine atom in place of a hydrogen, which makes the name methyl chloride, or chloromethane. The next two have ethane as the base alkane, can you see the two carbon atoms? Bromoethane (or ethyl bromide) and iodoethane (or ethyl iodide) are the molecules shown here. 'Iodo-' or 'iodide' are the terms we use when iodine replaces a hydrogen atom; 'bromo-' or 'bromide' are used for bromine, and so on. The colour used in the diagrams are randomly chosen. Chlorine is not green, and bromine is not red. We could have chosen any colour, but often use green to represent a halogen in molecular models.

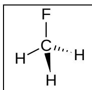
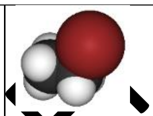
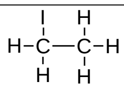
 a. fluoromethane or methyl fluoride	 b. chloromethane or methyl chloride	 c. Bromoethane or ethyl bromide https://commons.wikimedia.org/wiki/Category:Bromoethane#/media/File:Bromoethane-3D-vdW.png	 d. iodoethane or ethyl iodide

Figure 9.14 Different representations of alkyl halides

The halogens can also replace hydrogen atoms in alkenes or alkynes, for example forming bromoethene or chloropropyne. This is an example where the molecule contains two functional groups, namely a halogen and a double or triple bond.

Uses of the alkyl halides

Haloalkanes are often used in fire extinguishers, aerosol propellants, in refrigeration and to make foamed plastic. Dry cleaning processes also use haloalkanes, for example chloroethane, instead of water as the solvent. You can watch a video about the hydrocarbons and alkyl halides here: https://youtu.be/ZhwQoZR5rzE

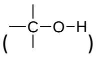
Substituents

When a halogen, or any other atom or group of atoms, replace a hydrogen in a carbon chain, it is called a **substituent**. When the substituent contains carbon and hydrogen atoms, for example $-CH_3$, it is called an alkyl group. For example, methyl- is used when there is only one carbon, ethyl- for two carbons, propyl- for three carbons, and so on.

DEFINITION: Substituent

A substituent is an atom or group of atoms bonded to a carbon chain. This can be an inorganic atom (e.g. halogen) or an alkyl group that is shorter than the main group.

Alcohols

The next homologous series is the alcohols. An alcohol is any organic compound where there is a **hydroxyl functional group** () bound to a carbon atom. The general formula for a simple alcohol is $\mathbf{C_nH_{2n+1}OH}$. The simplest and most commonly used alcohols are methanol and ethanol. Different representations of them are shown in Figure 9.15.

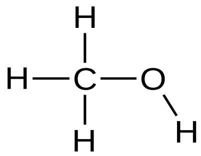
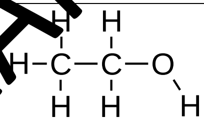
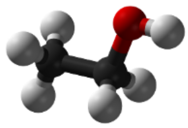
Name of alcohol	Condensed structural formula	Structural formula	Three-dimensional models
Methanol	CH_3OH		 https://upload.wikimedia.org/wikipedia/commons/thumb/8/86/Methanol-3D-vdW.png/110px-Methanol-3D-vdW.png
Ethanol	CH_3CH_2OH		 https://commons.wikimedia.org/wiki/File:Ethanol-3D-balls.png

Figure 9.15 Methanol and ethanol

Uses of alcohols

The alcohols have many uses. The most well-known is probably the use of ethanol in alcoholic beverages. Ethanol is also widely used as solvent in industrial processes, medicinal products and perfumes, and as an antiseptic. Methanol is mixed with ethanol in methylated spirits. Alcohols are toxic, for example if methanol is ingested it forms formic acid or formate salts, which damage the central nervous system and can cause blindness, coma or death. Ethanol is the only alcohol that is not toxic when ingested in low concentrations, but in higher concentrations it is also toxic.

Carbonyl groups

A carbonyl group contains a carbon atom to which an oxygen atom is linked through a double bond. The carbon atoms are further bonded to two other groups, which we indicate using R and R' as shown in Figure 9.16. R and R' represent other groups of atoms, for example alkyl groups, or hydrogen. Different groups of compounds are formed depending on which groups of atoms attached to each side of the carbonyl group. For example, an aldehyde is formed when an alkyl group and a hydrogen attach to form a -CHO group as you can see in Figure 9.16b. A ketone is formed when carbon chains attach to both sides of the carbonyl group as in Figure 9.16c and a carboxylic acid forms when a hydroxyl group (-OH) and a carbon chain attach to the carbonyl group as in Figure 9.16d. Aldehydes, ketones and carboxylic acids form three different homologous series. We will discuss aldehydes and ketones in this unit, and carboxylic acids in the next unit.

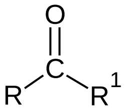
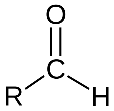
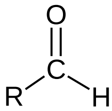
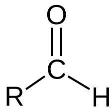
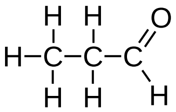
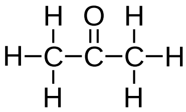
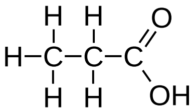
 a. general formula for a carbonyl group	 b. general formula for aldehydes	 c. general formula for ketones	 d. general formula for carboxylic acids
	Propanal 	Propanone 	Propanoic acid 

Figure 9.16 The carbonyl group is present in aldehydes, ketones and carboxylic acids

You can watch a video on different functional groups here: https://youtu.be/Uf5CvbOIP98 or here: https://youtu.be/hlXc_eEtBHA. Some compounds like ethers and amines mentioned in this video we won't be covering in this course, so you don't have to remember them.

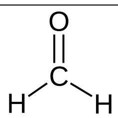
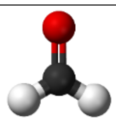
Commented [MR1]: This and the ones before are mindset videos with South Africans – may be better

Commented [MR2]: I don't particularly like this one.

Aldehydes and ketones

Aldehydes and ketones both have carbonyl groups in their structures. Aldehydes can be recognised by a –C=O group at the **end** of the molecule. A hydrogen atom replaced the R' group to form –C(O)H as you can see in Figure 9.17a and b. Ketones have the –C=O group in the **middle** of the molecule with carbon chains on both sides. Figure 9.17c and d shows two ketone molecules.

The first aldehyde is called methanal or formaldehyde and has only one carbon atom. The second aldehyde has two carbon atoms and is called ethanal or acetaldehyde. You will notice that all aldehyde names end in **–al**'. The next two aldehydes in this homologous series are propanal with three carbon atoms, and butanal with four carbon atoms.

Name of aldehyde or ketone	**Structural formula**	**Semi-structural formula**	**Ball-and-stick model**
a. methanal or formaldehyde		CH_2O or HCHO	 https://commons.wikimedia.org/wiki/File:Formaldehyde-3D-balls.png

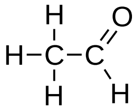
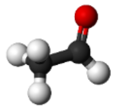
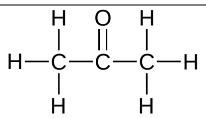
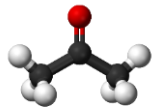
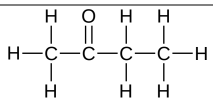
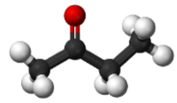
b. ethanal or acetaldehyde		CH_3CHO	 https://commons.wikimedia.org/wiki/File:Acetaldehyde-3D-balls.png
c. propanone or acetone		$CH_3C(O)CH_3$	 https://commons.wikimedia.org/wiki/File:Acetone-3D-balls-B.png
d. butanone		$CH_3C(O)CH_2CH_3$	 https://en.wikipedia.org/wiki/File:Butanone-3D-balls.png

Figure 9.17 Aldehydes and ketones

The first ketone is propanone, a molecule with three carbon atoms (Figure 9.17c). Why do you think methanone and ethanone, molecules with one and two carbon atoms respectively, cannot exist? Ketones have the carbonyl group in the middle of the molecule, and not at the end like the aldehydes. This means that at least one carbon atom needs to be connected on each side of the carbonyl group. Ketones must therefore have at least three carbon atoms in their structures. A ketone with four carbon atoms is called butanone. You will again notice that all the ketones have the same name ending, namely **–one**.

A video lesson on aldehydes and ketones can be found here: https://youtu.be/c2mSz1m3BSo.

Uses of aldehydes and ketones

Methanal is also known as formaldehyde and is used in embalming, tanning, germicides, insecticides and fungicides for plants. However, its main use is in the plastics industry where it is reacted it with phenol to form Bakelite. This was one of the first plastics made and was used to make jewellery, billiard balls, buttons and plastic cups and saucers. You can watch a video on the history of Bakelite here: https://youtu.be/vU5PCfz6b24.

Ketones are generally used as solvents and catalysts in the chemical industry, as stabilizers in paints or in perfumes. Propanone is more commonly known as acetone, a well-known solvent.

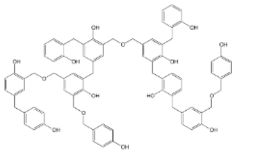
 A Bakelite radio https://en.wikipedia.org/wiki/Bakelite	 A Bakelite telephone https://en.wikipedia.org/wiki/Bakelite	 https://commons.wikimedia.org/wiki/File:Phenolic-resin-base-catalysis-5.png

Figure 9.18 uses of aldehydes and ketones

Activity 9.7 Recognising functional groups

The box below contains the structures of aldehydes, ketones, alcohols and alkyl halides. Identify the homologous series to which each of these compounds belong.

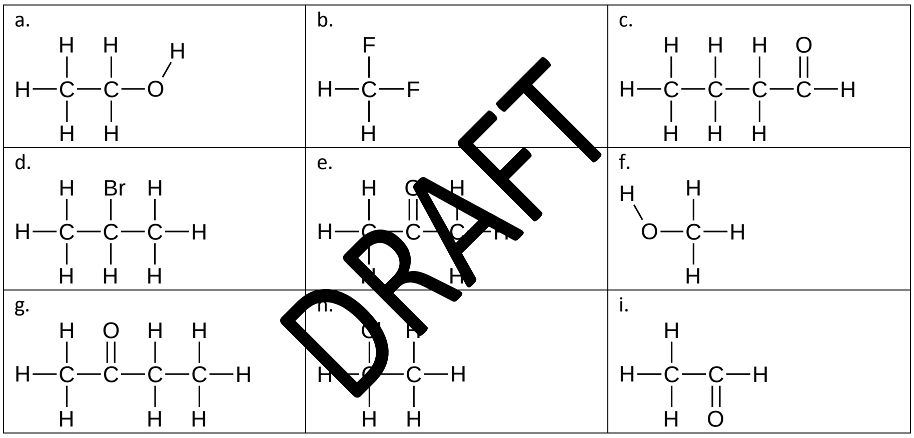
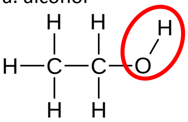
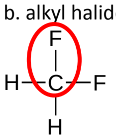
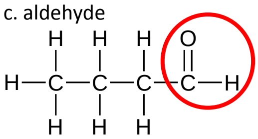
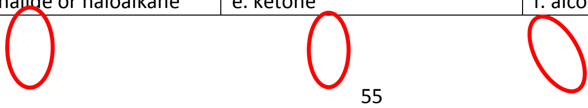


Figure 9.19 Recognising functional groups

Activity 9.7 Guided reflection

Compare your answers to the ones below (Figure 9.20). Which ones did you struggle with the most? The easiest way to identify the homologous series is to look for the atoms that are not carbon and hydrogen. If there aren't any, you have the hydrocarbons, but there weren't any in this activity. If you find a halogen, you have the alkyl halides. If you find an oxygen and it is linked to a hydrogen to form -OH, you have the alcohols. If you find the C=O at the end of the molecule as -CHO, you have the aldehydes and if the C=O is in the middle of the molecule you have the ketones. If you have a carbonyl (C=O) and an alcohol (-OH) linked to the same carbon, it is a carboxylic acid. This activity did not have any carboxylic acids.

a. alcohol 	b. alkyl halide or haloalkane 	c. aldehyde 
d. alkyl halide or haloalkane	e. ketone	f. alcohol



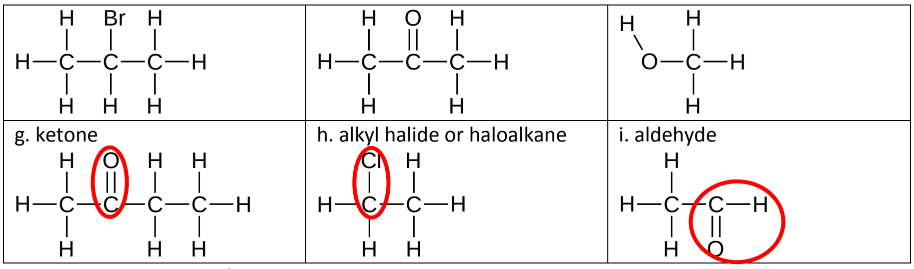


Figure 9.20 Recognising functional groups answers

The IUPAC naming system

The International Union of Pure and Applied Chemistry (IUPAC) is an international organisation responsible for the standardisation of names, symbols and units in chemistry. The organic chemistry nomenclature (naming system) is governed by IUPAC to ensure consistency across the world. Since there are so many organic compounds, and since new ones are made almost every day, it is important to have a system where each compound has a unique name. In this section you will learn how to name some of the basic organic structures.

The following steps will guide you in naming organic molecules.
For a video explanation for these steps you can watch the following: https://youtu.be/nsggKuxlYHw or https://youtu.be/U7wavimfNFE.

Step 1:

Identify the functional group in the compound. This will determine the suffix or ending of the name. Table 9.5 lists the functional groups and suffixes you will use in this course.

Table 9.5 Functional groups and suffixes

Functional group	Suffix
Alkane	-ane
Alkene	-ene
Alkyne	-yne
Alcohol	-ol
Aldehyde	-al
Ketone	-one
Carboxylic acid	-oic acid
Ester	-oate

Step 2:

Find the longest continuous carbon chain that contains the functional group. This does not have to be a straight chain. Count the number of carbon atoms in the chain. This will determine the prefix of the compound's name. Table 9.6 lists the first eight prefixes. You have already come across the first four, for example ethanol has two carbon atoms, and butanone has four carbon atoms. In this course you will only learn the first eight names, but many more exist, for example nine carbon compounds use **non-** and ten carbon compounds use **deca-**.

Table 9.6 number of carbon atoms and prefixes

Number of carbon atoms	Prefix
1	Meth-
2	Eth-

3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-

Step 3:

Number the carbons in the longest continuous carbon chain. If there is a functional group, the numbering is such that the functional group takes the lowest possible number. To do this, you start with the number at the end that is closest to the functional group. This is called the parent chain.

Step 4:

Look for any side chains. These are alkyl substituents, in other words carbon-hydrogen groups, for example ethane that have replaced a hydrogen in the main chain. The name of the group will change to end in -yl, for example methyl for one carbon, ethyl for a two-carbon group, propyl for a three-carbon group, etc.

When more than one of the same side chains are present, use the following prefixes to indicate the number: di- for two groups, tri- for three groups, and tetra- for four groups, for example dimethyl, triethyl, or tetrapropyl.

The side chains must be listed, in alphabetical order, before the name of the parent chain.

Step 5:

For the alkyl halides, the halogen group is treated the same as a hydrocarbon side chain. The name of the halogen is changed to end in -o, for example chloro- or bromo- and listed alphabetically. Di-, tri- and tetra- are used to indicated multiple halogen substituents, but these are not considered when placed in alphabetical order.

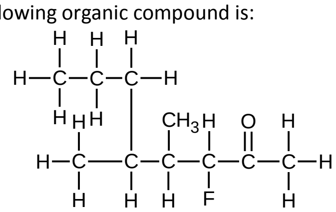
Step 6:

Combine the different parts of the name into a single word in the following order:

1. hydrocarbon side chains and halogens in alphabetical order, ignoring prefixes
2. prefix of the main chain
3. suffix of the main chain according to the functional group and position on the longest chain.

Example:

Find the IUPAC name of the following organic compound is:



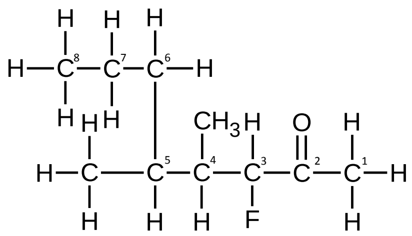
Step 1: Identify the functional group

There is a carbonyl group (C=O) present with carbon atoms on both sides, so it is a KETONE. It will have a **-one** suffix.

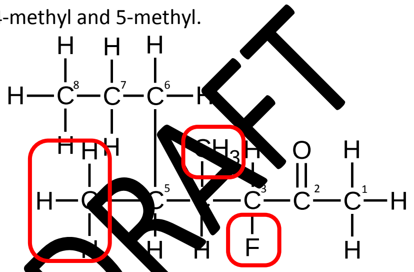
Step 2: Find the longest continuous carbon chain

There are eight carbons in the longest chain, so the name will have an **octan-** prefix. Note that the horizontal six-carbon chain is not the longest one.

Step 3: Find the position of the functional group.
The ketone group is on carbon 2 (or carbon 7, but 2 is lower), so the name of the parent chain is **octan-2-one**.



Step 4: Identify the side chains
Once the parent chain is numbered, the numbering cannot be altered. In this case, numbering starts from the 'right hand side'. There is a fluorine atom on carbon 3 and a methyl group on carbons 4 and 5. The groups are 3-fluoro, 4-methyl and 5-methyl.



Step 5: Combine identical side chains
When there are two or more identical side chains, they are combined and di-, tri- or tetra- is used. The groups are also arranged alphabetically according to the name of the side chain. The side chains for this molecule become 3-fluoro-4,5-dimethyl-. Note that the 'di-'prefix is not considered when placed in alphabetical order.

Step 6: Combine the parts to form the IUPAC name:
The final name is 3-fluoro-4,5-dimethyloctan-2-one. Note how we use the dash (-), the comma (,) and no spaces. This is quite important and should be done correctly. The following names are INCORRECT:
3,fluoro-4,5-dimethyloctan-2-one
3-fluoro-4-5-dimethyloctan-2-one
3-fluoro-4,5-dimethyl octan-2-one

For a detailed explanation of naming hydrocarbons, watch the following video. The video is quite long, but worth watching: https://youtu.be/HqLIk_F78yU.

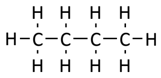
We are now going to do examples for each homologous series. You need to practice naming compounds regularly so for each homologous series there is an exercise.

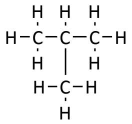
Naming alkanes

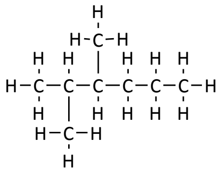
The suffix for an alkane is -ane. Alkanes contain only single bonds, and only carbon and hydrogen atoms.

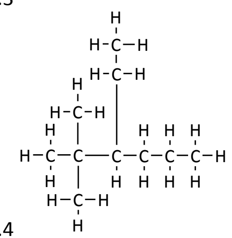
Activity 9.7 Naming alkanes

1. Provide the IUPAC names for each of the following molecules.

1.1 

1.2 

1.3 

1.4 

2. For each of the compounds above, write the condensed structural formula.

Activity 9.7 Guided reflection

Compare your answers to the ones below. The IUPAC name and condensed structural formula are provided.

1.1 Butane, $CH_3CH_2CH_2CH_3$

1.2 2-methylpropane, $CH_3CH(CH_3)CH_3$

1.3 2,3-dimethylhexane, $CH_3CH(CH_3)CH(CH_3)CH_2CH_2CH_2CH_3$ (notice that there are several ways of finding the chain of 6 for hexane but you need to choose the one where the branch begins on the lowest number)

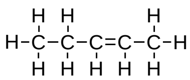
1.4 3-ethyl-2,2-dimethylhexane, $CH_3C(CH_3)_2CH(C_2H_5)CH_2CH_2CH_3$

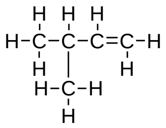
Naming alkenes

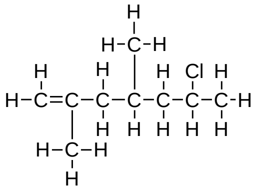
The suffix for an alkene is -ene. Alkenes contain at least one double bond, and only carbon and hydrogen atoms.

Activity 9.8 Naming alkenes

1. Provide the IUPAC names for each of the following molecules.

1.1 

1.2


1.3


2. Write down the molecular formulae for each of the structures above.

Activity 9.8 Guided reflection

Compare your answers to the names below.

2.1 pent-2-ene, C_5H_{10}
Remember to use the smallest number. You sometimes need to start counting from the other side of the molecule.

2.2 3-methylbut-1-ene, C_5H_{10}

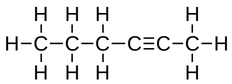
2.3 6-chloro-2,4-dimethylhept-1-ene, $C_9H_{17}Cl$. Did you manage to name this one correctly? Make sure you get the numbers, commas and dashes correctly written down.

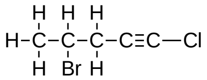
Naming alkynes

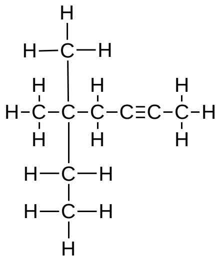
The suffix for an alkyne is -yne. Alkynes contain at least one triple bond, and only carbon and hydrogen atoms.

Activity 9.9 Naming alkynes

1. Provide the IUPAC names for each of the following molecules.

1.1


1.2


1.3


2. What is the general formula for the alkynes?

Activity 9.9 Guided reflection

1.1 hex-3-yne

1.2 4-bromo-1-chloropent-1-yne

1.3 5,5-dimethylhept-2-yne

2. C_nH_{2n-2}

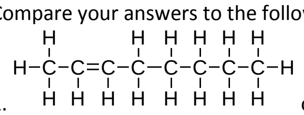
Activity 9.10 Drawing hydrocarbons

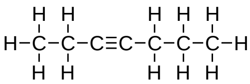
Draw the structural formulae for the following hydrocarbons:

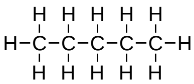
1. oct-2-ene
2. hept-3-yne
3. pentane
4. propyne
5. 2-methylbut-1-ene
6. 3-methylhexane
7. 4-ethyl-4-methylhept-2-yne

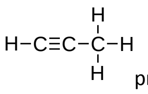
Activity 9.10 Guided reflection

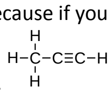
Compare your answers to the following:

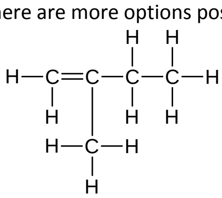
1.  oct-2-ene

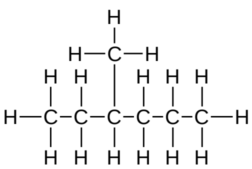
2.  hept-3-yne

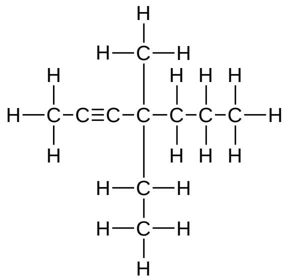
3.  pentane

4.  propyne.

Note that propyne does not require a number (e.g. prop-1-yne) because if you drew the double bond between carbon 2 and carbon 3, you get the same molecule.  is the same molecule as the one above, you just started numbering from the other side. You only include a number where there are more options possible.

5.  2-methylbut-1-ene

6.  3-methylhexane

7.  4-ethyl-4-methylhept-2-yne

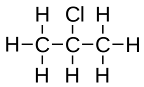
A video lesson on branched hydrocarbons and alkyl halides can be found here:
https://youtu.be/ZhwQoZR5rzE

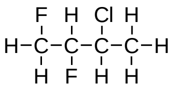
Naming alkyl halides

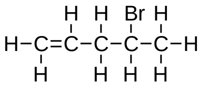
The halogens that often bond with carbon are fluorine, bromine, chlorine and iodine. Remember that the halogen atom name changes to, for example, chloro or bromo, when it is a substituent in an organic molecule.

Activity 9.11 Naming alkyl halides

1. Give the IUPAC name for each of the following molecules:

1.1 

1.2 

1.3 

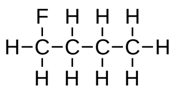
2. Write the structural formula for each of the following compounds:

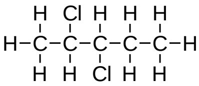
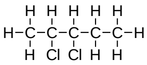
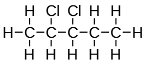
2.1 1-fluorobutane
2.2 2,3-dichloropentane
2.3 3-iodohex-2-ene

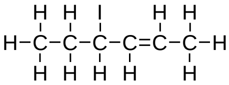
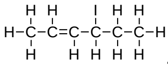
Activity 9.11 Guided reflection

Compare your answers to the ones below.

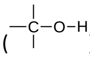
1.1 2-chloropropane
1.2 1,2-dibromo-3-chlorobutane. The prefix di- is not considered, the halogens are ranked alphabetically, bromo- before chloro-.
1.3 4-bromopent-1-ene. Note that the double bond takes preference over the halogen and the molecule will be 4-bromopent-1-ene and not 2-bromopent-4-ene

2.1 

2.2  Note that the following representations  and  are also accepted.

2.3  This structure can also be written as .

Naming alcohols

The alcohols contain the hydroxyl functional group () and the name ends in or -ol. If there are more than one hydroxyl group, the name becomes a diol (for two hydroxyl groups) or a triol (for three hydroxyl groups). Just like for the alkyl groups, you use numbers to indicate where the -OH groups are. For example, hexan-2,3-diol has a hydroxy group on carbons 2 and 3.

You can watch a video about naming alcohols here: https://youtu.be/nQ7QSV4JRSs.

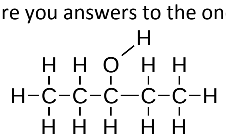
Activity 9.12 Naming alcohols

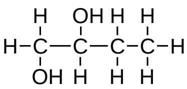
1. Draw the structural formulae for each of the following molecules:
1.1 pentan-3-ol
1.2 butane-1,2-diol
1.3 2-chloropropan-1-ol
1.4 4-ethyloctan-2,5-diol

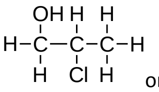
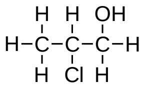
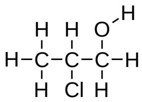
2. Draw the condensed structural formulae for each of the compounds above.
3. An unknown alkane diol has three carbon atoms. Draw all the possible structures for this compound and include the IUPAC names.

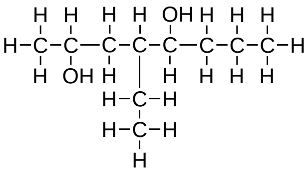
Activity 9.12 Guided reflection

Compare you answers to the ones below.

1.1 

1.2  . Note that you can draw the two hydroxyl groups on the same, or on opposite sides of the molecule.

1.3  or  Note that you can show the bond between O and H as shown in this representation  . It is sometimes acceptable to not show this, due to space constraints, but in general rather add this bond in.



1.4

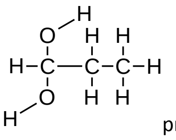
2.1 $CH_3CH_2CH(OH)CH_2CH_3$

2.2 $CH_2(OH)CH(OH)CH_2CH_3$

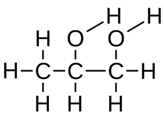
2.3.1 $CH_2(OH)CH(Cl)CH_3$

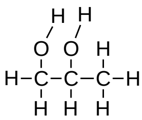
2.4 $CH_3CH(OH)CH_2CH(C_2H_5)CH(OH)CH_2CH_2CH_3$

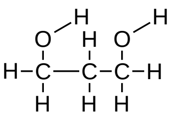
3. A diol has two hydroxyl groups. These can be on the same carbon atom, or on different ones. An alkane with three carbon atoms in propane. Note also that when we name diols, propane keeps the 'e' at the end. Usually it will be left out, for example with propan-2-ol (no 'e' after the 'n'), but this does not happen for diols and triols. The reason has to do with pronunciation.



propane-1,1-diol (and not propan-1,1-diol)



propane-1,2-diol. Note that this structure  is the same.



propane-1,3-diol

Naming aldehydes and ketones

Remember that aldehydes do not need a number to indicate the position of the aldehyde group as this will always be on the first carbon. When other substituents are present, the carbonyl group takes the preference, in other words, you always start your numbering with the carbonyl group on carbon 1. Aldehydes end in -al (or -anal) and ketones end in -one. You can watch a video on naming aldehydes and ketones here: https://youtu.be/c2mSz1m3BSo.

Activity 9.13 Naming aldehydes and ketones

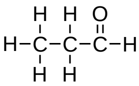
1. Write the structural formula for the following compounds.

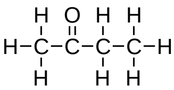
1.1 propanone

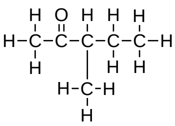
1.2 octanal

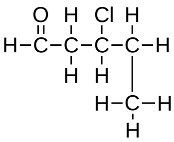
1.3 2-fluoro-4-methylhexan-3-one

2. Write the IUPAC names for the following compounds.

2.1 

2.2 

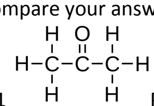
2.3 

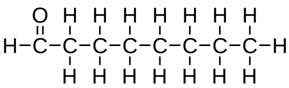
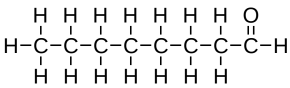
2.4 

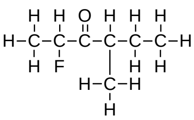
3. Write the molecular formula for each of the compounds in question 1.
4. Write the structural formula for the functional group of aldehydes.
5. What is the name of the functional groups of ketones?

Activity 9.13 Guided reflection

Compare your answers to the ones below.

1.1  Propanone has three carbon atoms with the carbonyl group on the 'middle' carbon.

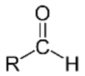
1.2  or  Here you should have eight carbon atoms with the aldehyde group at the end of the molecule, either the left on the right hand side.

1.3 

2.1 propanal. Don't forget to count the carbon that is part of the carbonyl group (C=O).

2.2 butanone. Note that this is not butan-2-one as the carbonyl group can only be in one position to be a ketone. Butan-3-one is an incorrect name, as it is the same as butan-2-one, just numbered from the other side.
2.3 3-methylpentan-2-one. The carbonyl group determines the numbering, and in this example, is on carbon 2.
2.4 3-chloropentanal. Did you have butanal here? Count carefully when looking for the longest continuous carbon chain as it may go 'around the corner'.

3.1 C_3H_6O
3.2 $C_8H_{16}O$
3.3 $C_6H_{13}OF$

4. 

5. Carbonyl group

You will find an organic chemistry revision exercise after unit 10.

Key learning points from unit 9

Now that you have completed the study of unit 9, you should be able to:

1. Describe how organic chemistry is integral to our daily lives.
2. Explain the difference between organic and inorganic compounds.
3. Define the terms hydrocarbon, functional group, homologous series.
4. Identify the homologous series to which a molecule belongs.
5. Identify the following functional groups in molecules: alkanes, alkenes, alkynes, alkyl halides, alcohols, aldehydes and ketones.
6. For each functional group you should be able to:
   - name the functional group.
   - make a model of a molecule containing a functional group using commercial kits, or sweets/food and toothpicks.
   - draw a projection formula for molecules with maximum two carbon atoms.
   - draw the structural, semi-structural, condensed and molecular formulae.
7. Provide IUPAC names for organic molecules with a maximum of eight carbon atoms and three functional groups. Substituents will only be alkyl or halide groups.
8. Identify everyday examples and uses for organic molecules.

Unit 10: Carboxylic acids and esters

Outcomes

By the end of this unit, you should be able to

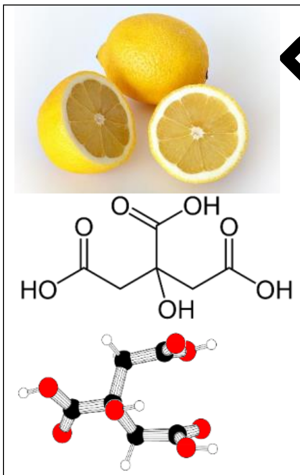
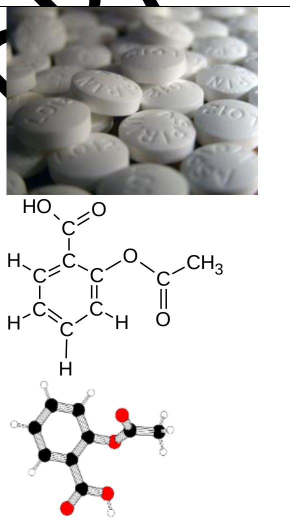
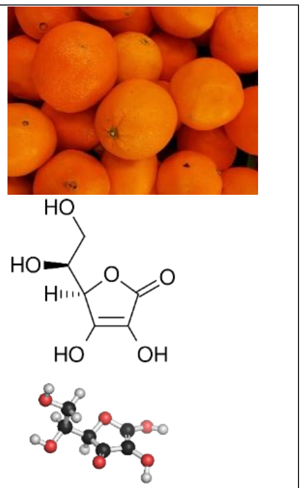
- draw the structural, semi-structural, condensed and molecular formulae of unbranched carboxylic acids up to eight carbon atoms.
- provide the IUPAC names for carboxylic acids up to eight carbon atoms and with a maximum of one substituent, either an alkyl group or halide.
- describe the carboxylic acids as weak acids in terms of the Lowry-Bronsted model.
- use IUPAC names and structural formulae to describe the reaction of a carboxylic acid with an alcohol to form an ester, e.g. methanol + ethanoic acid → methyl ethanoate + water.
- provide the IUPAC name for esters up to eight carbon atoms.
- identify commercial applications of esters.

Introduction to unit 10

In the previous unit you learned about organic molecules, how to name them and why they are important to us. In this unit we are going to learn about two additional homologous series, namely the carboxylic acids and the esters. This unit continues on from unit 9, so if you need some revision of the introduction to organic chemistry and IUPAC naming of compounds, you can watch the following two videos: https://youtu.be/YAjpntbbQ2w and https://youtu.be/nsggKuxlYHw.

Carboxylic acids

Carboxylic acid are organic molecules found in many household products. Figure 10.1 shows a few such products. You may notice that these products all have a sour taste. This is because they contain an organic acid. Lemons contain citric acid, aspirin is used for headaches and as a blood thinning agent and contains acetyl salicylic acid, while oranges contain vitamin C, which is ascorbic acid.

 a) lemons contain citric acid, an organic carboxylic acid https://upload.wikimedia.org/wikipedia/commons/thumb/e/	 b) aspirin is acetyl salicylic acid and used for headaches https://commons.wikimedia.org/wiki/Category:Aspirin#/med	 c) oranges contain ascorbic acid also known as vitamin C https://commons.wikimedia.org/wiki/Category:Oranges_as_food#/media/File:Navelate_2.jpg

e4/Lemon.jpg/1024px-Lemon.jpg https://commons.wikimedia.org/wiki/Citric_acid#/media/File:Zitronens%C3%A4ure_-_Citric_acid.svg https://commons.wikimedia.org/wiki/Citric_acid#/media/File:Citric-acid-3d.png	ia/File:%D8%A7%D9%84%D8%A3%D8%B3%D8%A8%D8%B1%D9%8A%D9%86.jpg https://commons.wikimedia.org/wiki/File:Aspirin-rod.png	https://commons.wikimedia.org/wiki/Category:Ascorbic_acid#/media/File:Ascorbic_acid_structure.svg https://commons.wikimedia.org/wiki/Category:Ascorbic_acid#/media/File:Ascorbic_acid_3D_model.png

Figure 10.1 Various representation of carboxylic acids in everyday life

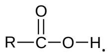
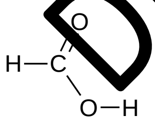
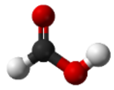
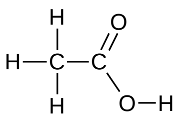
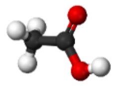
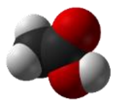
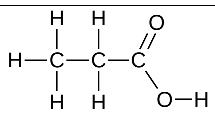
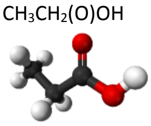
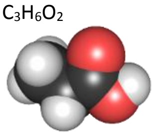
Carboxylic acids are organic molecules which contain a carbonyl group (-C=O), just like aldehydes and ketones, but have a hydroxyl group (-OH) bonded on the one side of the carbonyl group to form a -COOH group. Note that we should in fact write this as -C(O)OH, but it is quite acceptable to write -COOH. This functional group is always found at the one end of the molecule. The functional group is called a **carboxyl** group and has the general formula .

Figure 10.2 gives the first four carboxylic acids, namely methanoic acid, ethanoic acid, propanoic acid and butanoic acid. This follows the same naming pattern for the prefixes as in unit 9. You will notice that the names all end in **-anoic acid**. The general formula for the carboxylic acids is $\mathbf{C_nH_{2n+1}COOH}$, for example propanoic acid has three carbon atoms so n=3. The molecular formula for propanoic acid will be $C_3H_{2(3)+1}COOH$ which is C_3H_7COOH.

Name	Structural formula	Semi-structural formula and ball and stick model	Molecular formula and space filling model
Methanoic acid also known as formic acid		HC(O)OH  https://commons.wikimedia.org/wiki/File:Formic-acid-3D-balls-A.png	CH_2O_2  https://commons.wikimedia.org/wiki/File:Formic-acid-3D-vdW.png
Ethanoic acid also known as acetic acid		$CH_3C(O)OH$  https://commons.wikimedia.org/wiki/File:Acetic-acid-3D-balls.png	$C_2H_4O_2$  https://commons.wikimedia.org/wiki/File:Acetic-acid-CRC-GED-3D-vdW-B.png
Propanoic acid also known as propionic acid		$CH_3CH_2(O)OH$ 	$C_3H_6O_2$ 

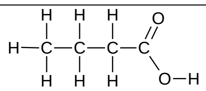
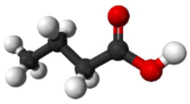
		https://commons.wikimedia.org/wiki/File:Propionic-acid-3D-balls.png	https://commons.wikimedia.org/wiki/File:Propionic_acid_spheres.png
Butanoic acid also known as butyric acid		$CH_3CH_2CH_2C(O)OH$  https://commons.wikimedia.org/wiki/Category:Butyric_acid#/media/File:Butyric-acid-3D-balls.png	$C_4H_8O_2$  https://commons.wikimedia.org/wiki/File:Butyric-acid-3D-vdW.png

Figure 10.2: The first four carboxylic acids

You have already seen that carboxylic acids are common in nature. Methanoic acid, also known as formic acid is found in insect stings. Ethanoic acid, or acetic acid, is vinegar, and benzoic acid, a more complex carboxylic acid, is used as a food preservative. Next time you look at a food label, see if you can find some of these on the label.

Activity 10.1

https://sites.google.com/a/g.coppellisd.com/per-1-hailey-fatima/food-label
Look at the ingredients on the food label above. How many organic acids can you see?

Guided reflection

There are many *organic compounds* such as sugar, maltodextrin, and lactose but these are not acids. the organic acids mentioned are lactic acid and citric acid.

You can have a look at an introduction to carboxylic acids in the following video: https://youtu.be/xheOq0XZ-so. You don't have to remember the movement of electrons in between the bonds in a carboxylic acid that is shown in this video.

You may have noticed that some of the molecules we discussed in unit 9 have alternative names, for example acetone instead of propanone or acetaldehyde instead of ethanal. This is because some of these compounds were known and named long before a standardised naming system was introduced. For example, methanoic acid has one carbon atom and it is also known as formic acid. This comes from the Latin word 'formica' which means 'ant'. Ants secrete formic acid as a defence mechanism. The aldehyde with one carbon (methanal) is known as formaldehyde and the formyl group is -CHO. Since these terms have been part of everyday language long ago, they are still being used today. However, since many more compounds have been made since, a standardised naming system for all organic molecules was needed and today we use the IUPAC naming system.

IUPAC naming for carboxylic acids

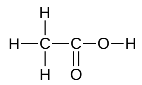
Carboxylic acids are characterised by having **a carboxyl group**, which has the formula **–COOH** at the end of the molecule. The name of a carboxylic acid ends in **-oic acid**.

Activity 10.2

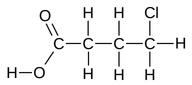
1. Draw structural formula for each of the following compounds:
1.1 pentanoic acid
1.2 4-ethyl-6-methyloctanoic acid

2. Provide the IUPAC names for each of the following molecules.

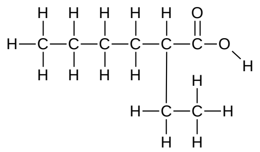
2.1



2.2



2.3



2.4 $CH_3CH_2CH(CH_3)CH_2CH_2CH_2CH_2COOH$

3. What is the common or household name for each of the following:

3.1 ethanoic acid

3.2 methanoic acid

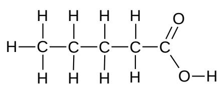
3.3 butanoic acid

3.4 propanoic acid

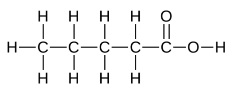
Guided reflection

Compare your answers to the ones below.

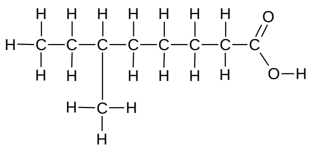
1.1 Pentanoic acid looks like this



Remember that pentanoic acid has five carbon atoms. You can also draw the structure like this with the COOH group in a straight line.



1.2 Remember to start counting from the carboxyl group's side. Then you will have the ethyl group on carbon 4 and the methyl group on carbon 6.



2.1 The name is ethanoic acid. Do not write acetic acid when they ask for the IUPAC name, as acetic acid is the household or common name used for this compound, ethanoic acid is the IUPAC name.

2.1 The name is 4-chlorobutanoic acid. Do not get confused when the molecule is drawn 'upside down' or swopped 'left-to-right'. It is still the same molecule.

2.3 The name is 2-ethylhexanoic acid

2.4 The name is 6-methyloctanoic acid

Carboxylic acids should not be too difficult to name. You just need to look out for the -COOH group and then start numbering from there as the carbon of the carboxyl group takes the lowest number.

3.1 acetic acid
3.2 formic acid
3.3 butyric acid
3.4 propionic acid

Carboxylic acids as weak acids

Carboxylic acids are acids and have similar properties to inorganic acids. Acids have a sour taste and change the colour of acid-base indicators, for example litmus paper from blue to red. In the next activity you are going to verify these properties of a carboxylic acid.

Activity 10.3

For this activity you will need vinegar, baking soda and an indicator, like blue litmus paper or universal indicator. You will also need a drinking glass and a teaspoon. If you do not have litmus paper or universal indicator you can make your own indicator from red cabbage or petunia flowers. This video will show you how it is done: https://youtu.be/lEdfYik5iX8

Pour a small amount of vinegar in a glass. You need about 30 ml. Dip a teaspoon in the vinegar and touch a piece of blue litmus paper (or your own home made indicator as shown in the video above) with the wet tip of the teaspoon. What do you observe? You can watch a short video clip of this experiment here: https://youtu.be/X6hS1qiqzeQ

Add 10 drops of universal indicator to the left-over vinegar in your glass. What do you observe? Add a teaspoon of baking soda, stir until the bubbles have stopped and observe. Write down what you observe. Are there any changes in the colour of the solution? Not yet, then add another teaspoon of baking soda, stir, and observe. Keep on adding baking soda, one teaspoon at a time, until no more bubbles form and the colour of the solution changed.

Guided Reflection

Litmus is an acid-base indicator that turns red in an acid solution and blue in a base. It is usually supplied as paper strips, see Figure 10.3. You will learn more about acids and bases in unit 17. When you dipped the litmus paper into the vinegar it turned red. This shows that vinegar is an acid.

Litmus paper, universal indicator and red cabbage juice change colour in vinegar.

Figure 10.3 Vinegar in various indicators

When a base like baking soda is added to an acid, the acidic properties are cancelled by the base. When you have universal indicator present, the colour will change from red to yellow and then to green as you add more base. If you used red cabbage juice, the colour should also be pink-red in the beginning. The bubbles that are formed in this reaction are carbon dioxide gas. The balanced equation for the reaction between vinegar (ethanoic acid) and baking soda (sodium hydrogen carbonate) is shown below:

$CH_3COOH_{(aq)} + NaHCO_{3(s)} \rightarrow CH_3COO^-_{(aq)} + Na^+_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Vinegar is a carboxylic acid. It is acidic as it changed the colour of an acid-base indicator. It also reacts with baking soda to form carbon dioxide gas and water and changes the colour of the indicator as the solution changes from acidic to basic.

 *a. vinegar with universal indicator shows a red colour*	 *b. vinegar solution with two teaspoons baking soda shows a yellow colour and bubbles are present*	 *c. vinegar solution with four teaspoons baking soda shows a green colour and no more bubbles are being formed*

Figure 10.4 The reaction between vinegar and baking soda

The reactions of carboxylic acids with water

The vinegar you used in Activity 10.2 is not pure acetic acid, it is a dilute solution of about 5% acetic acid in water. When any carboxylic acid dissolves in water it reacts with the water to form the carboxylate ion. A hydrogen ion (proton) is donated to a water molecule to form the hydronium ion (H_3O^+). The presence of hydronium ions in water give the solution its acidic properties. The equation below shows how acetic acid reacts with water. Can you see that you start with CH_3COOH which then gives an H^+ to water, which becomes H_3O^+. The CH_3COOH becomes the acetate ion, CH_3COO^-.

$$CH_3COOH_{(l)} + H_2O_{(l)} \leftrightarrows CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

acetic acid + water ⇆ acetate ion + hydronium ion

Acetic acid is described as a weak acid. Weak acids do not release hydrogen ions easily. We say it ionises partially. Organic acids are all weak acids. You will learn more about weak and strong acids in unit 17.

Activity 10.4

A few drops of lemon juice are dripped onto blue litmus paper. Lemon juice contains the organic substance citric acid.

1.1 *What colour change do you expect on the litmus paper?*
1.2 *Explain your prediction in question 1.1.*
1.3 *Which ion is responsible for your prediction in question 1.3?*

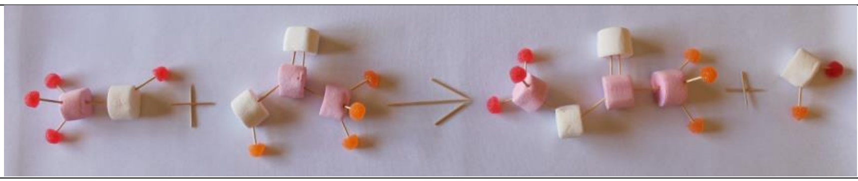
2. Propanoic acid is dissolved in water. Write a balanced chemical equation to show how propanoic acid dissolves in water.

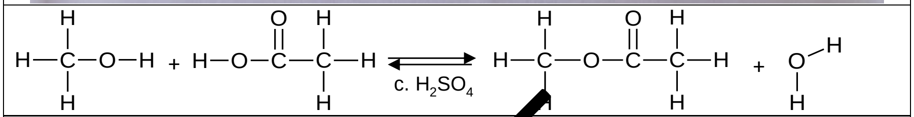
Guided reflection

1.1 The blue litmus paper will turn red. All acids will turn litmus red and citric acid is an acid.
1.2 Citric acid has acidic properties and will change blue litmus red.
1.3 The hydronium ion or H_3O^+ is responsible for the colour change.
2. $CH_3CH_2COOH_{(l)} + H_2O_{(l)} \leftrightarrows CH_3CH_2COO^-_{(aq)} + H_3O^+_{(aq)}$

Esters

When an alcohol reacts with a carboxylic acid, an ester is formed. The name of the functional group is an **ester** group and the structural formula is $R-COO-R^1$. This is a new homologous series with the general formula $C_nH_{2n}O_2$. In the reaction between the alcohol and the ester, a molecule of water is removed from the two compounds and a new bond is formed between what remains of the alcohol and the carboxylic acid. This reaction is called **esterification**. A catalyst is required in this reaction, in this case it must be an inorganic acid like H_2SO_4. The reaction between ethanoic acid and methanol is shown in Figure 10.5 built with models, in symbolic form, and in words.





methanol + ethanoic acid $\rightleftharpoons$ (c. H_2SO_4) methyl ethanoate + water

Figure 10.5: The reaction between an alcohol and a carboxylic acid forms an ester and water

Activity 10.5

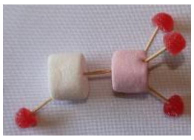
For this activity you are going to need a commercial molecular model set, or sweets like marshmallows, jelly tots and toothpicks. Just like in unit 9, we are going to build structure for organic molecules, but this time we want to investigation how esters are formed.

Step 1: Build a methanol molecule using one pink marshmallow (representing a carbon atom), one white marshmallow (representing an oxygen atom) and three jelly tots (representing hydrogen atoms). Use Figure 10.5 as a guide.
Step 2: Build ethanoic acid using two carbon atoms, two oxygen atoms and four hydrogen atoms. Again use Figure 10.5 as a guide.
Step 3: Remove the hydrogen from the methanol which is attached to the oxygen. Remove the oxygen and its hydrogen together from the ethanoic acid.
Step 4: Connect the remainder of the methanol, now called a methyl group (CH_3), to the remainder of the ethanoic acid, now called an ethanoate group. This forms the ester methyl ethanoate. The bond that connects these two parts is called and **ester bond**.
Step 5: You are left with a hydrogen atom and a hydroxyl group (-OH). Connect these two to form a water molecule.
Step 6: Compare your products to the pictures in Figure 10.5.

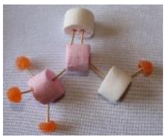
Guided reflection

Esters are formed when an alcohol and a carboxylic acid reacts in the presence of a catalyst, like concentrated sulfuric acid (H_2SO_4). You will have noticed that a water molecule was formed as a by-product of this reaction. A by product is another product that we are not interested in at this point. The sulfuric acid acts as a catalyst in this reaction. You will learn more about catalysts in unit 14. It lowers the activation energy and makes the reaction take place faster but does not itself take part in the reaction. It also acts as a dehydrating agent because it removes water from the reaction.

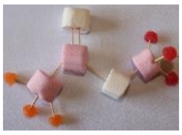
Step 1: Your methanol molecule should look like this:



Step 2: Your ethanoic acid molecule should look like this:



Step 4: Your ester should look like this:



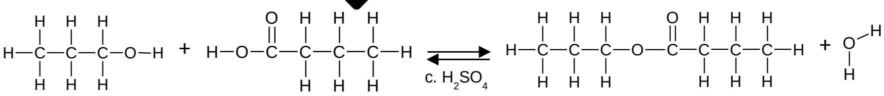
Remember that the toothpicks represent bonds. This does not mean that a chemical bond is a 'stick' that you can pull out and place somewhere else. A chemical bond is an area where you are more likely to find the electrons or an area of high electron density, like you learned about in unit 6.

Have you noticed that the name of the ester you have made is a combination of the names of the two reactants and ends in **-oate**. We are now going to look at how esters are named.

Naming esters

The name of an ester is derived from the names of the alcohol and carboxylic acid that was used to form it. Let's look at the following example:

Propanol is reacting with butanoic acid to form an ester according to the following reaction:



The ester that is formed is called **propyl butanoate**. The first part of the name of this ester is derived from the alcohol, propanol, so **propyl** is used. The second part of the name is derived from the carboxylic acid butanoic acid and is changed to -oate, namely **butanoate**. The name is therefore propyl butanoate. Note that the name for an ester is written in two separate words. The name of the ester in Figure 10.6 below is methyl propanoate. If you draw an imaginary line between the O and the carbonyl group, then the oxygen side is methyl because it has one carbon atom, and the carbonyl side is the propanoate as it has three carbon atoms. This also means that this ester was formed from methanol and propanoic acid.

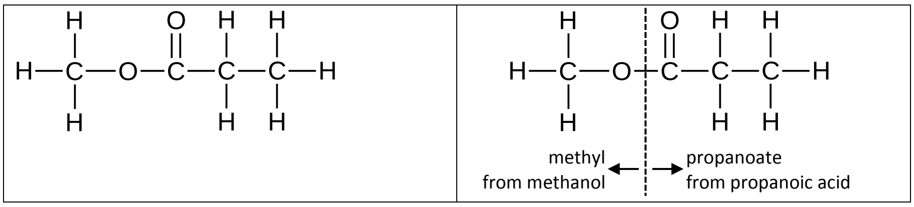
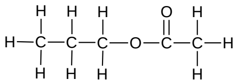


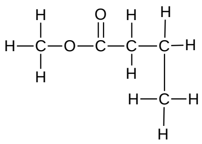
Figure 10.6: An example of an ester

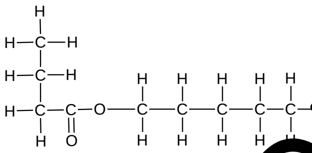
Here is an activity where you can practice naming esters. Look carefully as the molecules can be swopped around so that the alkyl is on the right-hand side, and not the left-hand side as in the example above.

Activity 10.6 Naming esters

1. Provide the IUPAC names for each of the following esters:

1.1 

1.2 

1.3 

2. Identify the alcohol and carboxylic acid that were used to prepare each of the following esters.
2.1 butyl methanoate
2.2 propyl pentanoate
2.3 ethyl ethanoate
2.4 methyl hexanoate

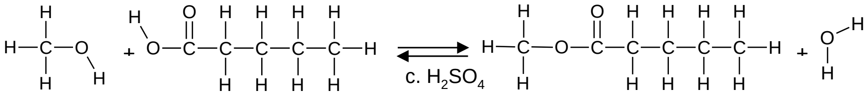
3. Write a balanced chemical equation for the reaction between the following two compounds. Include all the products, and name all the compounds in the reaction.
3.1 methanol and pentanoic acid
3.2 methanoic acid and ethanol

Guided reflection

1.1 propyl ethanoate.
1.2 methyl butanoate. Don't get confused by the chain that is drawn 'around a corner'. It is still a continuous carbon chain so you count it as if it is drawn straight.
1.3 pentyl butanoate. Don't let the fact that the molecule is drawn 'upside down' confuse you. Just find the carbonyl group, identify the number of carbon atoms on the oxygen side (five in this case) and write 'pentyl'. Then find the number of carbon atoms on the carbonyl side, including the carbonyl atom (four in this case) and finish the name by writing butanoate.
2.1 butanol and methanoic acid
2.2 propanol and pentanoic acid

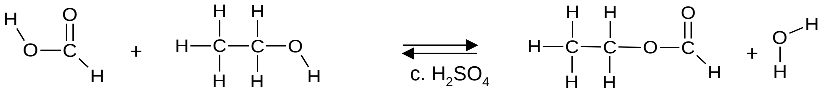
2.3 ethanol and ethanoic acid
2.4 methanol and hexanoic acid
This question is easier to do. Did you also find it easier? The first part of the ester name is derived from the alcohol, and the second part from the carboxylic acid.

3.1 Compare your answer with the equation below. Remember to add in the sulfuric acid, and don't forget the water molecule.



methanol + pentanoic acid → methyl pentanoate + water

3.2 Did you get this one right? The acid was given first, don't let that put you off. You can draw the acid and the alcohol in any order on the left-hand side of the arrow.



methanoic acid + ethanol → ethyl methanoate + water

Properties and uses of esters

Most esters have a characteristic smell and are used extensively in the fragrance and flavours industry. For example, pentyl ethanoate smells like banana and is used as a flavour for banana sweets. Figure 10.7 shows a large number of smells associated with esters. Esters are also used in cosmetics and beauty products because they usually have a fruity smell, making them useful as artificial flavourants and scents. They are used in nail varnish remover and model plane glue, as solvents for non-water-soluble compounds like oils and resins, and as plasticisers because they make products more flexible.
If you click on the link, you can download a bigger picture of figure 10.7

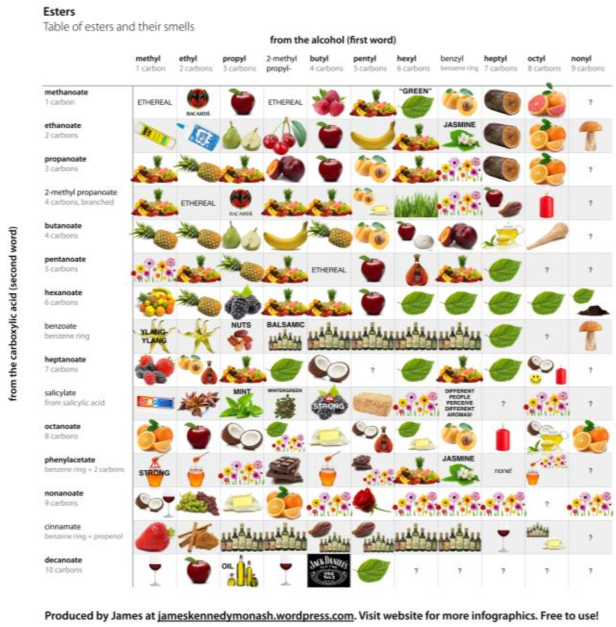


Figure 10.7 Esters and their smells (https://jameskennedymonash.files.wordpress.com/2013/12/table-of-esters-and-their-smells-v2.pdf) [This is free to use, but I am not sure if the quality will be acceptable. Perhaps set it to be a page big, or put the link and remove the figure if the quality is not good.]

We are now going to investigate how to make esters in the laboratory. If you are able to get hold of some of the chemicals and equipment, do try this experiment yourself.

Activity 10.6 Making esters

As we have learned, esters are made by reacting an alcohol with a carboxylic acid in the presence of concentrated sulfuric acid. Watch the following video to see how the experiment is done: https://youtu.be/AILt5HAYJTs

The table below represents a flow diagram for the steps in this reaction. Fill in the missing words in the table and answer the questions that follow. You may need to watch the video again to find the answers.

Flow diagram for the esterification reaction between ethanol and ethanoic acid.	
	1. Pour 2 ml _____________ to the test tube.

	2. Pour about 2 ml _____________ acid to the test tube.
	3. Add 2-3 drops of _____________ _____________ _____________. This acts as a _____________.
	4. Place the beaker into a _____________ _____________ bath. Heat until the mixture just _____________ _____ _________. Remove the test tube to let it _____________ _____________ so that the content just does not boil, but stays hot.
	5. Separate the product from the reaction mixture by adding it to a solution of _____________ _____________.
	6. _____________ start to form when the reaction mixture is added. A _____________ is observed. The product smells like _____________ _____________.

Answer the following questions:

1. What is the general equation to produce an ester.
2. What inorganic acid is needed in an esterification reaction?
3. What special safety precaution is needed when you smell a product?
4. What does ethanoic acid smell like?
5. In the second experiment solid carboxylic acid was used. What was the name of the carboxylic acid?
6. What did the product in the second experiment smell like?

Activity 10.6 Guided reflection

Compare your flow diagram to the one below.

	This is an esterification reaction between ethanol and ethanoic acid. 1. Pour 2 ml **ethanoic acid** to the test tube.
	2. Pour about 2 ml **ethanol** acid to the test tube.
	3. Add 2-3 drops of **concentrated sulfuric acid**. This acts as a **catalyst**.
	4. Place the beaker into a **hot water** bath. Heat until the mixture just **starts to boil**. Remove the test tube to let it **cool down** so that the content just does not boil, but stays hot.
	5. Separate the product from the reaction mixture by adding it to a solution of **sodium carbonate**.
	6. **Bubbles** start to form when the reaction mixture is added. A **smell** is observed. The product smells like **paint thinners**.

1. Alcohol + carboxylic acid → ester + water
2. Concentrated sulfuric acid
3. Do not smell substances by putting your nose directly into a test tube. Instead, waft the smell towards your nose.
4. It smells like vinegar.
5. Salicylic acid
6. Wintergreen

You have reached the end of unit 10. Below is a final assessment for units 9 and 10 together. If you need to revise some nomenclature, you can watch a video on naming oxygen containing compounds here. This is a bit of revision of unit 9, as well as carboxylic acids and esters from this chapter: https://youtu.be/1xdXRqR9W0Q

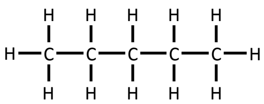
Summary Assessment for Units 9 and 10

1. Complete the table below.

Name	Name of homologous series	Structural formula	Semi-structural formula	Condensed structural formula	Molecular formula	Name and structural formula of the functional group
methanol						
propanoic acid						
ethanal						
1-chloropropane						
butane						
hex-2-ene						
heptan-3-one						
ethyl butanoate						

2. Provide IUPAC names for the compound below.

2.1



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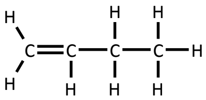

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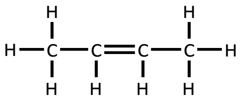

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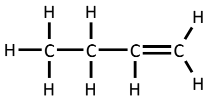

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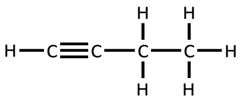

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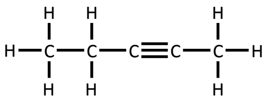

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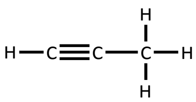

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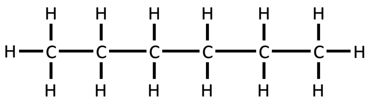

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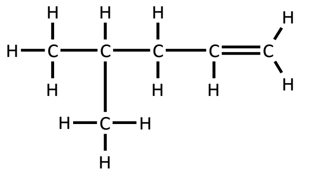

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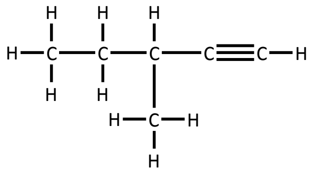

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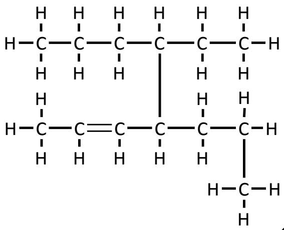

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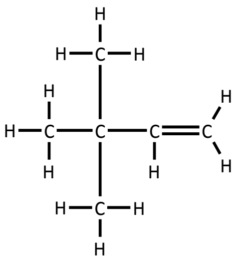

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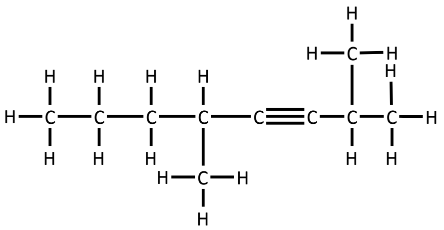

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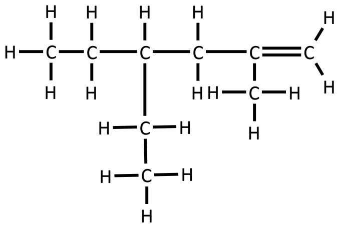

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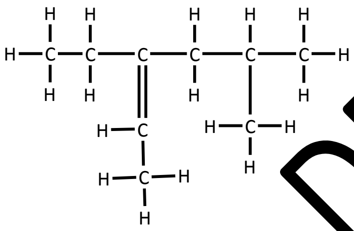

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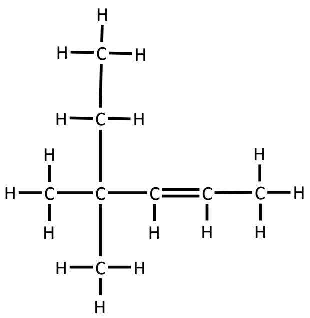

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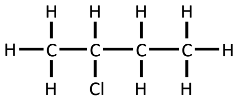

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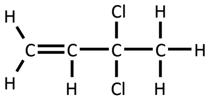

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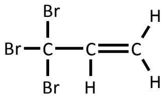

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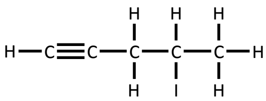

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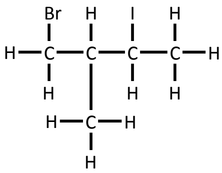
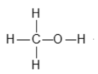
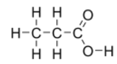
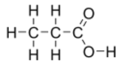
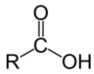
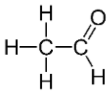
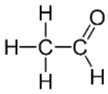
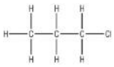
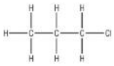
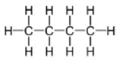
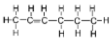
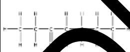


Answer Key for Summary Assessment

Compare your answers to the answers in the table below. Which ones did you find most challenging to draw?

IUPAC name	Name of homologous series	Structural formula	Semi-structural formula	Condensed structural formula	Molecular formula	Name and structural formula of the functional group
methanol	alcohol	CH_3OH		CH_3OH	CH_4O	ROH

IUPAC name	Name of homologous series	Structural formula	Semi-structural formula	Condensed structural formula	Molecular formula	Name and structural formula of the functional group
propanoic acid	Acids			$CH_3CH_2C(O)OH$	$C_3H_6O_2$	
ethanal	Aldehydes			CH_3CHO	CH_4O	R-C = O | H
1-chloropropane	Haloalkanes			CH_3CH_2Cl	C_2H_5Cl	RX
butane	Alkanes		$CH_3 – CH_2 – CH_2 – CH_3$	$CH_3(CH_2)_2CH_3$	C_4H_{10}	R
hex-2-ene	Alkenes		CH_3C[illegible]CH_2CH_3	$CH_3(CH)_2(CH_2)CH_3$	C_6H_{10}	R
Heptan-3-one	Ketones		$CH_3CH_2COCH_2CH_2CH_2CH_3$	$CH_3CH_2CO(CH_2)_3CH_3$	$C_7H_{14}O$	ROR'
Ethyl butanoate	Esters		$CH_3CH_2CH_2COOCH_2CH_3$	$CH_3(CH_2)_2COOCH_2CH_3$	$C_6H_{12}O_2$	ROOR'

2. Compare your IUPAC names with the names below.
2.1 pentane
2.2 1-butene or but-1-ene
2.3 2–butene or but-2-ene
2.4 1-butene or but-1-ene
2.5 1- butyne or but-1-yne
2.6 2-pentyne or 2 pent-2-yne
2.7 ethyne
2.8 propyne
2.9 hexane
2.10 4-methyl pent-1-ene *
2.11 3-methyl pent-1-yne*
2.12 4,5 propylethyl oct-2-ene*
2.13 2,2 dimethyl but-1-ene*
2.14 2,5 dimethyl oct-3-yne*

2.15 2,4 methylethyl oct-3-yne*
2.16 3, 5 ethylmethyl hex-2-ene*
2.17 4,4 dimethyl hex-2-ene
2.18 butane
2.19 3,3 dichloro but-1-ene
2.20 3, 3, 3 trichloro but-1-ene
2.21 4 iodopent-1-yne
2.22 1, 2, bromomethylbutane