

ELECTRON INTERACTIONS (2): AN OVERVIEW OF CHEMICAL REACTIONS

Learning Outcomes and Assessment Standards

Learning Outcome 1: Practical scientific inquiry and problem-solving skills.

Assessment Standard 2:

Interpreting data to draw conclusions.

Learning Outcome 2: Constructing and applying scientific knowledge.

Assessment Standard 1:

Recalling and stating specified concepts.

Assessment Standard 2:

Indicating and explaining relationships.

Overview

In this lesson we will study chemical reactions. We will apply many of the basic concepts covered in Lessons 15 & 16, Lessons 19 & 20 (specifically energy relations) and Lessons 21 & 22 (kinetic theory of matter) to a study of when and how atoms and molecules react.

- Molecular and formula mass
- Composition of substances
- Concentration
- Ions and radicals
- Stoichiometric calculations
- Energy changes in reactions related to bond energy changes
 - Exothermic and endothermic reactions
 - Activation energy
 - Energy changes in reactions related to bond energy changes
- Acid-base reactions
- Oxidation and reduction: redox reactions
- Substitution, addition and elimination reactions

Lesson



Molecular and formula mass

If you analyse a substance in the laboratory and find that it contained 1,5 parts of iron to two parts of oxygen, what would that tell you? Or if you analysed a sample of an organic substance (one containing mainly carbon) and found that it contained three parts carbon to six parts hydrogen to one part oxygen, what would you know with certainty?

Let's take the iron oxide. We know we cannot have fractions of an atom (see Dalton's Laws under Stoichiometric Calculations below), so the minimum it could be is Fe_3O_4 , which is still in the ratio 1,5:2 but raised to integers. But what if a molecule of this iron oxide was actually Fe_6O_8 ? And what if the organic compound was actually $\text{C}_6\text{H}_{12}\text{O}_2$? Both of these are quite possible. The fact is that this first analysis doesn't give a complete answer.

The simplest formula, with the simplest whole-number proportions, is known as the **empirical formula** because it gives the formula that we would calculate from an empirical laboratory analysis in which we determine the simple proportions of



the elements in a substance. We must use other methods as well as the empirical analysis to find the correct **molecular mass**. In the case of our examples above, we would get Fe_3O_4 . As a matter of interest, in the case of this iron oxide the story does not stop there. A still more detailed analysis will show that Fe_3O_4 is actually a combination of two different oxides of iron: FeO and Fe_2O_3 in a 1:1 ratio! If one calculates the oxidation number of iron in the two oxides, we find that FeO is iron (II) oxide – or **ferrous** oxide – and Fe_2O_3 is iron (III) oxide – or **ferric** oxide. (Remember that the higher oxidation number is named **-ic** and the lower **-ous**.)

Molecular mass (also called molecular weight)

We also speak of the **relative molecular mass (or weight)** of a compound because it is the mass of a molecule of the compound relative to a scale on which the mass of an atom of the ^{12}C isotope taken as exactly 12 u (or **atomic mass units**).

Formula mass (also called formula weight)

This term is used when we want to be strictly correct because the term “molecule”, according to the strict definition, only applies to groups of bonded atoms that can exist on their own. A crystal of sodium chloride (see Lessons 17 & 18) is actually composed of many, many Na^+ and Cl^- ions in a 1:1 ratio. When we use the formula “ NaCl ” it is on the understanding that there is actually no such independent entity as a sodium chloride molecule as such. Therefore we speak of NaCl as the formula and 58,442 u (= 22,989 u + 35,453 u) as being its **relative formula mass**. But in truth, these days the terms “molecular mass”, “molecular weight”, “formula mass” and “formula weight” are used interchangeably.

We will use the term **molecular mass** here even though we realise that not all formulae represent molecules in the strictest sense of the word.

CALCULATING MOLECULAR MASS (mm) AND THUS MOLAR MASS (M)

Carbon monoxide (CO)

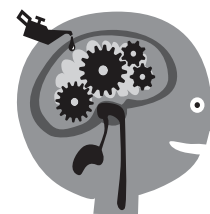
- CO is composed of one atom of carbon and one atom of oxygen.
- Atomic mass one carbon atom = 12,01 (from the Periodic Table)
Atomic mass of one oxygen atom = 16,00 (from the Periodic Table)
- Molecular mass for C = Atomic mass carbon + atomic mass oxygen

$$\begin{aligned} \text{mm}[\text{CO}] &= 12,01 \text{ u} + 16,00 \text{ u} \\ &= 28,01 \text{ u} \end{aligned}$$

- If the relative molecular mass of one molecule is 28,01 u then ...
- ... one mole of CO molecules has a mass of 28,01 g.
- \therefore Molar mass of CO is ... $M[\text{CO}] = 28,01 \text{ g}\cdot\text{mol}^{-1}$

Carbon dioxide, CO_2

- $\text{mm}[\text{CO}_2] = 12,01 + (2 \times 16,00) = 12,01 + 32,00 = 44,01 \text{ u}$
- $\therefore M[\text{CO}_2] = 44,01 \text{ g}\cdot\text{mol}^{-1}$



Water, H₂O

- $mm[\text{H}_2\text{O}] = (2 \times 1,008) + 16,00 = 2,016 + 16,00 = 18,016 \text{ u}$
- $\therefore M[\text{H}_2\text{O}] = 18,016 \text{ g}\cdot\text{mol}^{-1}$

Calcium hydroxide, [Ca(OH)₂]

- $mm[\text{Ca}(\text{OH})_2] = 40,08 + (2 \times 16,00) + (2 \times 1,008)$
 $= 40,08 + 32,00 + 2,016 = 74,096 \text{ u}$
- $\therefore M[\text{Ca}(\text{OH})_2] = 74,096 \text{ g}\cdot\text{mol}^{-1}$

Ammonium sulphate, (NH₄)₂SO₄

- $mm[(\text{NH}_4)_2\text{SO}_4] = (2 \times 14,01) + (8 \times 1,008) + 32,06 + (4 \times 16,00)$
 $= 28,02 + 8,064 + 32,06 + 64,00$
 $= 132,144 \text{ g}\cdot\text{mol}^{-1}$
- $\therefore M[(\text{NH}_4)_2\text{SO}_4] = 132,144 \text{ g}\cdot\text{mol}^{-1}$

COMPOSITION OF SUBSTANCES

After testing a substance, if we get an empirical formula of C₃H₆O₂ we need to perform other tests to see whether that is, indeed, the molecular formula as well. Let us assume that it is not. After further tests, we find that each molecule of the substance is actually C₆H₁₂O₄. Now the chemist's task is to find out the exact shape of the molecule and what atoms are joined where. For instance, is it the molecule C₅H₁₁COOH (hexanoic acid) or some other substance with the same ratio of C:H:O?

CONCENTRATION

When making up a solution, one dissolves:

- a known amount of a substance (the **solute**)
- in a known amount of **solvent** (e.g. water or alcohol).

It is important to know exactly how much of each we have. For example, in **solution A** we have 5,286 g of (NH₄)₂SO₄ in 250 cm³ of water; while in **solution B** there is 1,321 g of (NH₄)₂SO₄ in 500 cm³.

If we take exactly 100 cm³ of each solution and evaporate the solvent (water) we will be left with two different amounts of the solute, i.e. ammonium sulphate. What is important to know in each case is the ratio of the mass of solute to the volume of solvent. That ratio is what we call the **concentration**.

The standard unit of concentration is “**moles per dm³**”. If we know the molar mass of the **solute** (e.g. 132,144 g·mol⁻¹ for ammonium sulphate) and the volume of **solvent** it was dissolved in, we can calculate the concentration using the formula:

$$\text{concentration} = \frac{\text{moles}}{\text{volume}}$$

$$\text{or } c = \frac{n}{V}$$

The number of moles of solute, $n = \frac{\text{mass}}{\text{molar mass}} = \frac{m}{M}$

$$\therefore c = \frac{m}{M \times V}$$

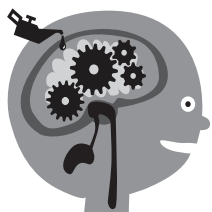
Calculations

- Concentration of solution A (5,286 g of (NH₄)₂SO₄ in 250 cm³ of water):

$$c = \frac{5,286}{132,144 \times 0,25} \text{ mol}\cdot\text{dm}^3$$

- Concentration of Solution B (1,321 g of (NH₄)₂SO₄ in 500 cm³ of water):

$$c = \frac{1,321}{132,144 \times 0,50} \text{ mol}\cdot\text{dm}^3$$



If we take 100 cm³ of each solution (or 0,1 dm³) and evaporate the solvent we will be left with 0,016 mol of (NH₄)₂SO₄ from solution A – or 2,11 g; and 0,020 mol of (NH₄)₂SO₄ from solution B – or 2,64 g.

STOICHIOMETRIC CALCULATIONS

In the early 1800s an English schoolteacher, John Dalton, put forward five laws of chemical combination that laid the basis for accurate chemical research. He was the first modern advocate of atomic theory, namely that all matter consisted of tiny particles (called atoms) and that every different element of matter was made up of identical atoms which were different from the atoms of every other element. Dalton's Laws are:

- All elements are composed of small particles called “atoms”.
- All atoms of the same element are identical.
- The atoms of a given element are different from those of any other element.
- Atoms cannot be created, divided into smaller particles nor destroyed. (Today we would add the words: “in the chemical process”.)
A chemical reaction simply changes the way atoms are grouped together.
- Atoms combine in simple ratios to form “compound atoms”. (Today we call these “molecules.”)

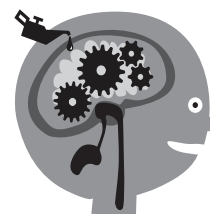
The most important conclusion from Dalton's Laws is that, in chemical reactions, substances will react in whole-number proportions. This gives us the basis of **stoichiometry** and what we know as **the laws of chemical combination**. (*Two of Dalton's “laws” are now known to be false, but at the time he formulated them there was no way of knowing this. We now know that not all atoms of an element are identical because a variable number of neutrons in the nucleus yields different isotopes. But the different isotopes of an element do not change the way atoms react chemically. The second law we now know to be false is the one about atoms being created or destroyed. Lessons 19 & 20 go into detail about nuclear fusion and fission. Dalton can hardly be blamed for not foreseeing developments in nuclear physics that only began 100 years later!*)

BONDING AND VALENCY; RADICAL IONS

As we know, in chemical reactions atoms react according to how many electrons they lose, gain or share. For example, if we describe an atom as having **a valency of one**, the atom might **either gain an electron** (as in the reaction between F or Cl with H, which loses an electron) **or lose an electron** (as in the reaction between Na or K reacting with F to produce NaF or KF). **The word valency applies to the number of electrons commonly lost or gained by single atoms.**

By knowing where the atom lies on the periodic table, we will know whether it is more likely to gain an electron and become a negative ion (atoms on the right-hand side of the table) or lose an electron and become a positive ion (atoms on the left-hand side of the table). Atoms can also share electrons and carbon, which lies in the middle of the periodic table, is the best example of a “sharer”. But we also get sharing when two atoms of the same kind form a molecule, e.g. F₂, Cl₂, O₂ or N₂.

In many reactions, a group of atoms with a charge (because of too many or too few electrons in the group) might take part in the reaction without undergoing any change. Examples of these are the acid radicals that we get when an acid dissolves in water. Sulphuric acid, H₂SO₄, and nitric acid, HNO₃, yield 2H⁺ + SO₄²⁻ and H⁺ and NO₃⁻ respectively. SO₄²⁻ and NO₃⁻ are the sulphate radical and the nitrate radical respectively.



(The table in Teacher Tips on page 342 gives some common valencies and the charge on various common radicals. It also gives instructions on the naming of some classes of compounds.)

ENERGY CHANGES IN REACTIONS RELATED TO BOND ENERGY CHANGES

The energy involved in chemical reactions is because bonds are being broken (where energy is given out) and new bonds are being made (where energy is taken in). The energy of a reaction is the overall amount of energy that is either given out during a reaction (a so-called **exothermic reaction**) or taken in during the reaction (a so-called **endothermic reaction**).

To find out more about the energy relationships in reactions, see the diagrams in Lessons 19 & 20 and the explanation of bond energies in Lessons 15 & 16.

Activation energy

It takes a certain amount of energy to get a reaction started. The energy that is put into a system to start the reaction is called the activation energy and we get this energy back once the reaction has started. In the diagrams below, the activation energy is represented by the shaded arrows. The activation energy is the energy required to get over the “energy hill”.

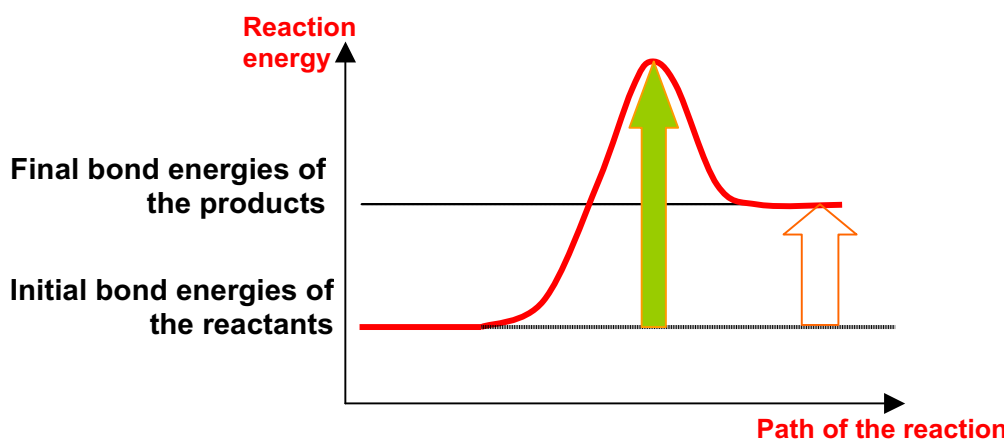
Endothermic reactions

When the total bond energies of the products are greater than the total for the reactants, it means that some energy had to be taken in from the environment. This energy is now locked up in the bonds of the products and can only be released when those bonds are broken again. When an endothermic reaction takes place in a container, it gets cooler because energy is being drawn into the container from the outside environment. The symbol used for energy of reactions is **H** (or “heat of reaction”).

For **endothermic** reactions, $\Delta H_{\text{reaction}} = H_{\text{final}} - H_{\text{initial}}$

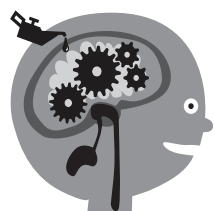
Since $H_{\text{final}} > H_{\text{initial}}$

$\therefore \Delta H_{\text{reaction}} > 0$ (i.e. positive)



Exothermic reactions

When the total bond energies of the reactants are greater than the total for the products, energy is passed into the environment. When exothermic reactions take place in a container, they heat up.

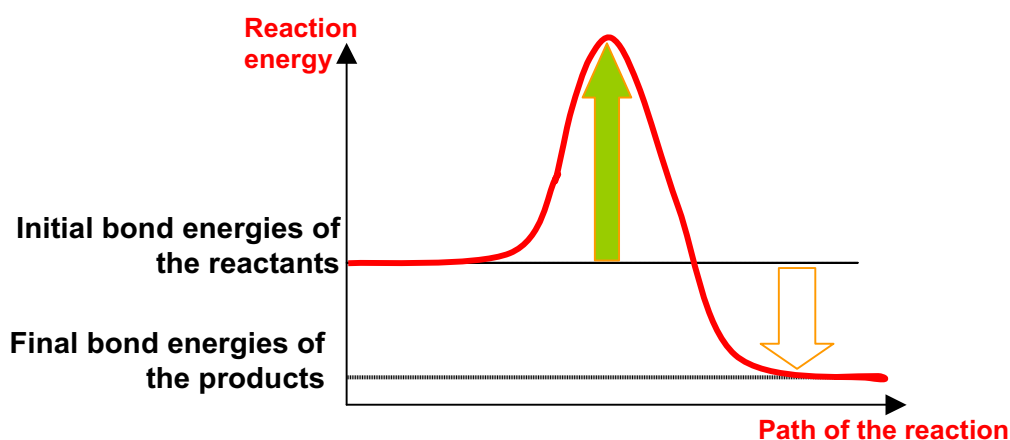


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For **exothermic** reactions, $\Delta H_{\text{reaction}} = H_{\text{final}} - H_{\text{initial}}$

Since $H_{\text{final}} < H_{\text{initial}}$

$\therefore \Delta H_{\text{reaction}} < 0$ (i.e. **negative**)



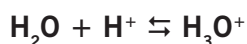
ACID-BASE REACTIONS

What is an acid?

The early definitions of acids and bases revolved around the sour taste of acids (as in acidic fruits and vegetables), their effect on natural vegetable colourings (like litmus or beetroot juice), the soapy feel of alkalis (they actually turn the oils in the skin into soap!) and their detergent qualities and their ability to break down oils and fats. In 1923, TM Lowry and JN Brønsted proposed that **acids** be regarded as substances that can **donate a proton** to another substance and **bases** are substances that **accept protons** (from other substances like acids!).



This is not the actual reaction because acid-base reactions occur in solution and the H^+ ion (a proton) cannot exist in that form in solution. The diameter of a hydrogen ion (i.e. a proton) is about 10^{-15} m. Other positive ions in solution are much bigger (about 10^{-10} m). Because the electric field around the much smaller proton is so intense (and positive), it attracts any molecule with unshared electrons, such as water. So for a proton in solution we get this reaction:



When **hydrogen chloride gas** dissolves in water we get this reaction taking place:

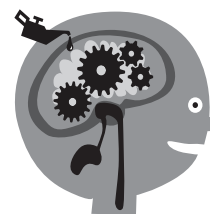


In this reaction the hydrogen chloride is the acid because it is donating a proton to the water molecule. Because we have an equilibrium, there are more hydronium ions (H_3O^+) than would normally be found in pure water. In acid reactions in solution with water, the hydronium ion is the proton donor. When a molecule like **HCl** separates into its ions we call the process **dissociation**.

When **ammonia gas** dissolves in water:



the water acts as a proton donor and the base, the ammonia gas, is the proton acceptor. By donating a proton, the water molecule produces an hydroxyl ion (OH^-) which is the powerful proton acceptor in acid-base reactions in solution.



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Water dissociates into the hydrogen and hydroxyl ions and the ammonia molecule and hydrogen ions associate to form an ammonium ion.

Acids and bases: strong and weak, dilute and concentrated

What makes an acid strong?

- For something to act as an acid, it must donate protons.
- The more protons a substance can donate, the stronger its acidity.
- In solutions, acids dissociate (or ionise) into their component ions.
- If one of those ions is a proton (H^+), the hydronium ion (H_3O^+) is formed.
- The more hydronium ions formed, the stronger the acid.
- So strong acids are substances that dissociate almost completely.

What makes an acid weak?

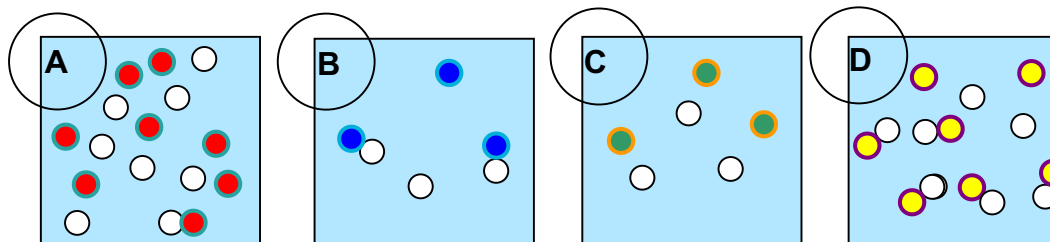
- And weak acids are substances whose molecules dissociate only slightly.
- Weak acids produce very few hydronium ions.

What makes bases strong or weak? (Use the same logic as for acids but change some of the terms!)

- ... base ... accepts protons.
- ... hydroxyl ions (OH^-).
- ... greater dissociation ... more hydroxyl ions ... stronger base.
- ... strong bases dissociate almost completely.

And now, what makes a solution (of acid or base) concentrated or dilute?

- Concentrated means a large number of molecules per unit volume.
- In solutions this means a lot of solute for not so much solvent.
- Dilute means a relatively small number of molecules per unit volume.
- In gases this means a few gas molecules relative to the volume of the container.



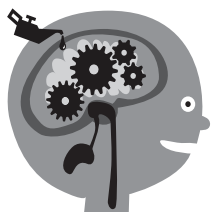
In the four diagrams representing four acids the white circles represent the hydrogen ion, H^+ , and the coloured circles represent acid radicals. Some of the acids in the diagrams have dissociated into ions.

Identify which of the diagrams **A-D** depict:

1. A **dilute, strong** acid?
2. A **concentrated, weak** acid?
3. A **dilute, weak** acid?
4. A **concentrated, strong** acid?

(The answers are at the end of this section on acids and bases.)

Types of acid: strong, medium and weak



The strength of an acid depends on what proportion of the acid molecules will dissociate when the acid is dissolved in water. We can define a number, K_a , for acids which is called the **dissociation constant**. If an acid is HA then $K_a = [\text{H}_3\text{O}^+][\text{A}^-]$ which is the product of the concentrations of the hydronium ions formed and the acid radical. For strong acids K_a is large and for weak acids K_a has a low value.

STRONG	<i>Sulphuric acid</i>	<i>Nitric acid</i>	<i>Hydrochloric acid</i>
	H_2SO_4	HNO_3	HCl
	$K_a \text{ approx } 10^3$	$K_a \text{ approx } 10^3$	$K_a \text{ approx } 10^3$
MEDIUM	<i>Phosphoric acid</i>	<i>Oxalic acid</i>	<i>Sulphurous acid</i>
	H_3PO_4	$(\text{COOH})_2$	H_2SO_3
	$K_a = 7 \times 10^3$	$K_a = 5,4 \times 10^{-2}$	$K_a = 1,7 \times 10^{-2}$
WEAK	<i>Carbonic acid</i>	<i>Hydrogen sulphide</i>	<i>Ammonium ion</i>
	H_2CO_3	H_2S	NH_4^+
	$K_a = 4,4 \times 10^{-7}$	$K_a = 1,3 \times 10^{-2}$	$K_a = 5,6 \times 10^{-10}$

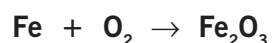
Look at the K_a value for the ammonium ion. $5,6 \times 10^{-10}$ is a very, very small number. What this indicates is that the chances of an ammonium ion dissociating into ammonia (NH_3) and a proton or hydrogen ion (H^+) are highly unlikely.

(ANSWERS TO DIAGRAMS: 1C, 2D, 3B, 4A)

OXIDATION AND REDUCTION: REDOX REACTIONS

What is oxidation?

Early chemists were aware of an important class of chemical reactions in which metals reacted with oxygen from the air. These were called “oxidation” reactions. The product of these reactions was an oxide, for instance “rust”:



The idea of oxidation numbers (see Lessons 17 & 18) was introduced to indicate the changes an atom would undergo in oxidation reactions.

During a chemical reaction ...

if the **oxidation number increases** an element is said to be **oxidised**.

If the **oxidation number decreases** the element is said to be **reduced**.

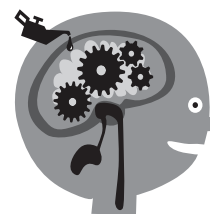
In the reaction above, iron (**Fe**) goes from **0** to **+3** i.e. the iron is **oxidised**;
oxygen (**O**) goes from **0** to **-2** i.e. it is **reduced**.

Although the word “**oxidation**” has remained with us, we now define the chemical process in terms of the movement of electrons rather than the presence of oxygen as a reactant.

OXIDATION is the ...

LOSS of electrons

If we use this definition of oxidation we see that it includes many other reactions besides those which involve oxygen. Another point to realise is that in a chemical reaction where electrons are LOST, something else must GAIN those electrons because we cannot have free electrons just “moving around”. So we define the other part of the process as:

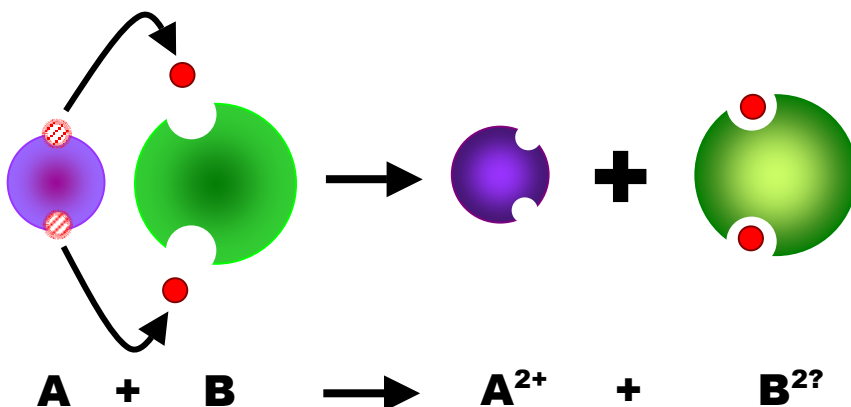


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REDUCTION is the ...

GAIN in the number of electrons

Oxidation and reduction reactions go hand-in-hand. If one substance **LOSES** electrons in a reaction (becomes **OXIDISED**) then something else must **GAIN** electrons (become **REDUCED**).



Reactions in which one substance is oxidised and another is reduced are known as oxidation – reduction reactions or **REDOX REACTIONS**.

In the picture we see that A loses two electrons and becomes A²⁺. In the same reaction B gains the two electrons lost by A and becomes B²⁻.

What is a half reaction?

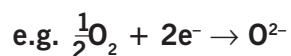
As we have seen, a redox reaction involves a loss of electrons by one substance and the gain of electrons by another. Even though these two processes must go together it is convenient to separate them out and look at the two halves of the reaction separately. We show the electrons in the half-reactions:



OXIDATION HALF-REACTION



REDUCTION HALF-REACTION



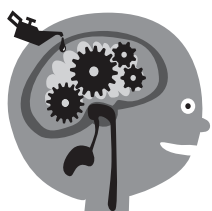
Balancing REDOX reactions

When one atom gives electrons and another takes them, we must ensure that **the number of electrons given is the same as the number of electrons taken**.

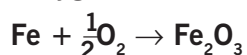
Making sure that the number of electrons is balanced is an excellent way of balancing the whole equation. Balancing the equation of a chemical reaction means that we make sure that:

- there is as much matter after the reaction has been completed as there was initially; and,
- if charged particles are involved, there is as much charge after as there was before.

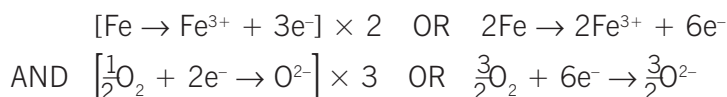
Using the iron (Fe) and oxygen (O) example again, we can see that **Fe + O → Fe₂O₃** is not balanced.



For a start, oxygen occurs naturally as a diatomic molecule so we denote one atom of oxygen like this:



But the number of atoms on both sides of a chemical equation must be equal so this is still unbalanced. By using the half reactions and balancing the electrons we get:



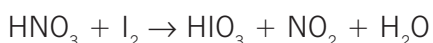
By adding the two equations we get **six electrons** on each side. By “cancelling the electrons” we understand that what is lost by the iron is gained by the oxygen. Because there is no such thing as “one and a half” oxygen molecules, we multiply through by two again to get:



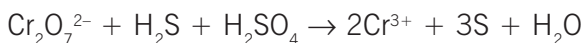
... we get a final equation which is: $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$

Examples

Look at the positions of nitrogen and sulphur on the periodic table. They are both in groups whose elements readily GAIN electrons. This makes them strong oxidising agents – i.e. they cause other atoms, ions and molecules to be oxidised (or lose electrons). Other well-known oxidising agents are the permanganate ion and the dichromate ion. Here are some examples of their redox reactions:



The nitrate ion is said to **oxidise** iodine to the iodate (IO_3^-) so the **NO_3^- ion** is the **oxidising agent**. The nitrate ion is reduced to nitric oxide, or **NO_2** . Check the change in oxidation numbers to verify this.



The dichromate ion **oxidises** the sulphide to pure sulphur. The **$\text{Cr}_2\text{O}_7^{2-}$ ion** is the **oxidising agent** and is itself reduced to **Cr^{3+}** .

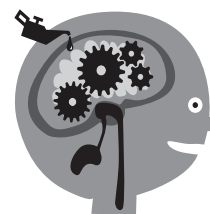
In order to balance a redox reaction we need at least FIVE facts:

- How likely is a reactant to gain or lose electrons?
- Which reactant will gain and which will lose electrons?
- How many electrons are gained or lost by a particular reactant?
- Are hydrogen ions involved (i.e. does the reaction take place in an acidic medium)?
- Is water involved in the reaction?

To balance equations for redox reactions we can use **tables of half-cell reactions** which give us the number of electrons gained and lost, the presence of H^+ ions and the involvement of water in the reaction. When water is used as a solvent to form solutions of aqueous ions it is not regarded as being *involved* in a reaction and is therefore not shown in the balanced equation.

ORGANIC CHEMISTRY

This is the branch of chemistry that deals with the chemical properties of one element: **carbon**. The fact that carbon atoms have four valence electrons and are much more likely to share electrons with other atoms than “give” or “take,” as in ionic reactions, is highly significant. For one thing it enables carbon to form long chains of carbon atoms. Other atoms can attach to the carbon on the chains and give the almost infinite number of possible carbon compounds many different



properties. Carbon chemistry is often called the “chemistry of life” because carbon, or organic, molecules form the basis of all processes in living organisms.

The four outer electrons of carbon arrange themselves in a **tetrahedral** shape (see VSEPR in Lessons 17 & 18) so the long chains of carbon actually follow a zigzag shape based on the angles of bonds around a tetrahedron.

If you can't picture a tetrahedron, take a small ball of plasticine or Prestik. Take four matches and stick them into the ball so that the angles between every pair of matches is the same.

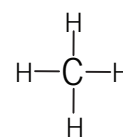
The shape you are holding is a tetrahedron!

Covalent bonding of carbon

Four carbon electrons can be shared with many other atoms that have single electrons to share. The carbon itself has four unshared electrons.

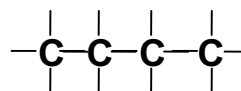


A carbon atom drawn like this shows four electrons available for bonding – but with nothing actually bonded yet.



And here we have a carbon atom whose four unshared electrons, have each formed a bond with the unshared electron from a hydrogen atom.

When carbons share electrons with each other they can form long chains.



Homologous series

Groups of organic molecules of the same type but which increase in size by one unit at a time are called **homologous series**. The simplest of these are the **hydrocarbons** whose molecules contain only carbon and hydrogen. And the simplest of the hydrocarbons are the **alkanes** in which we get only single bonds. The first four alkanes (methane, ethane, propane and butane) have unsystematic names, but thereafter they take their names from Greek numbers e.g. pentane (“penta” is Greek for “five”), hexane, heptane, octane, nonane, decane, etc.

Double bonds

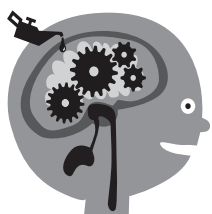
It is possible for two atoms that each have more than one unshared electron to make more than one bond. In other words, if two electrons from carbon shares with two electrons from another atom, e.g. oxygen, we get a **double bond**.

Chemical reactions are basically interactions involving electrons; double bonds are electron rich; and electrons in the double bond are quite exposed.

This makes double bonds highly reactive, i.e. **double bonds provide a ready site for chemical reaction.**

Functional groups

It is possible to form molecules that are based on carbon chains by replacing one or more hydrogen atoms atoms or groups of atoms that are called **functional groups**. Instead of homologous series of alkanes, for instance, we can now have series of substances like **alcohols, aldehydes, ketones, ethers, carboxylic acids,**



esters and **halo-alkanes** in which halogens use their single, unshared electrons to form single covalent bonds with carbon atoms. Another type of functional group is derived from the alkanes. When methane loses one hydrogen it becomes a **methyl radical** ($-\text{CH}_3$). The alkyl radicals also form an homologous series, i.e. **methyl-**, **ethyl-**, **propyl-** and **butyl-** groups, etc.

Functional group	Suffix	Functional group	Suffix
alcohol	-ol	aldehyde	-al
carboxylic acid	-oic acid	ketone	-one
ether	+ ether	ester	+ ester

Molecular and structural formulae

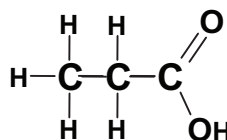
When we write formulae for organic molecules we can do so in three different ways. We can write the **molecular formula** in which we give the number of atoms of each different element in the molecule.

Example: $\text{C}_3\text{H}_6\text{O}_2$ [$mm = 3(12) + 6(1) + 2(16) = 74 \text{ u}$]

This formula is useful if we want to calculate the molecular mass but it gives no clue about the shape of the molecule. For this we need to know where the atoms are situated in the molecule. We also need to know what functional groups are present. The two kinds of structural formula give us this information:

Examples: $\text{CH}_3\text{CH}_2\text{COOH}$

Partial structural formula



Structural formula

SUBSTITUTION, ELIMINATION AND ADDITION REACTIONS

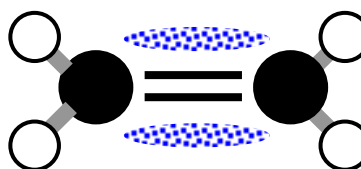
Reactions of organic compounds

Now that we know what organic molecules look like and where the electrons are to be found in the bonds we are in a better position to appreciate how they react.

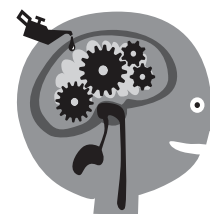
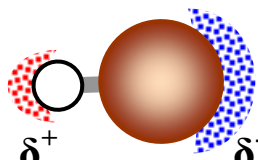
Addition reactions

This is a very important group of reactions that involves the double bond. Because of the greater number of electrons in the double bond and the fact that the bonding electrons do not all lie precisely between the double-bonded atoms, they are very easily involved in certain reactions.

Here is an example of a typical addition reaction between a molecule with two, double-bonded carbon atoms (the molecule is called "ethene" and it is a gas) and a molecule of hydrogen bromide.

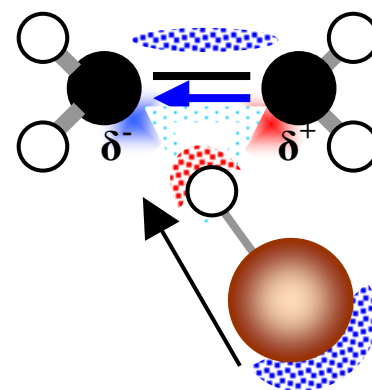


Because of the electronegativity difference (see Lessons 15 & 16) we know that HBr will be a polar covalent molecule. This means that the shared pair of electrons in the H-Br bond will separate leaving the bromide end more negative and the hydrogen end more positive.

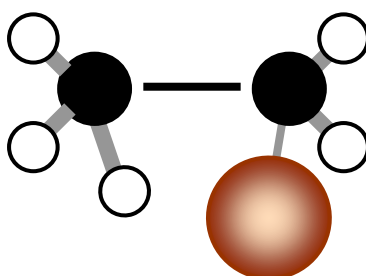




The positive end of the HBr molecule will “seek” the exposed electrons in the ethane double bond. (For this reason the HBr is called an “**electrophile**” because it is an “electron lover”.) When the positive end of the HBr approaches the double bond its electrons are pulled out of shape. This leaves one of the carbons slightly negative and the other one slightly positive. The positive and negative ends of the HBr will take advantage of the fact. The double bond is now broken and two new single bonds (one a C–H and the other a C–Br) are formed.

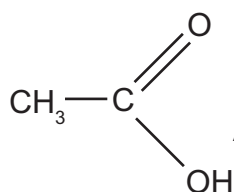


Once the approaching HBr has disturbed the distribution of electrons in the ethene the positive end (H⁺) can bond with the left-hand carbon and the Br⁻ part of the HBr will bond with the right-hand carbon. The double bond is broken but each carbon still have four bonds



CH₃CH₂Br
bromoethane

The example of HBr being added across the double bond was an example of electrophilic addition because the positive/electron-seeking part of the HBr molecule was active in breaking the double bond. It is also possible to get a different addition mechanism: nucleophilic attack. If a part of a molecule is exposed because of the electrons being drawn away, it is possible for a positive-seeking molecule to become attached to that point.

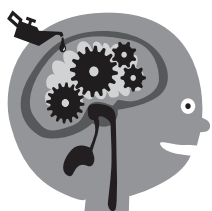


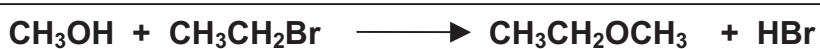
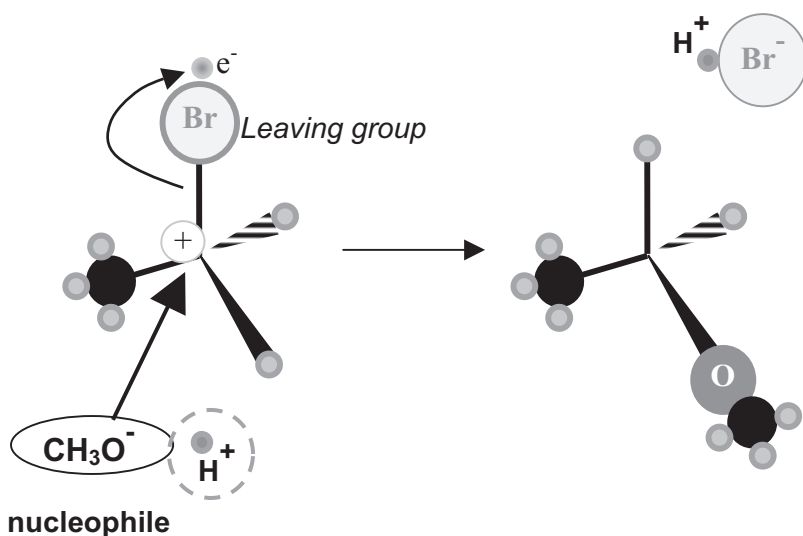
CH₃COOH
ethanoic acid
Also “acetic acid” / vinegar

The highly electronegative oxygens will draw electrons in the bonds away from the central carbon and leave it exposed and slightly positive. A molecule like water, with two unpaired electrons, can react with the ethanoic acid at this point.

Substitution reactions

In a **substitution reaction**, a functional group in a chemical compound is replaced by another group. Because of the high electronegativity of halogens compared to carbon, a carbo-halogen bond in a molecule will be **polarised** i.e. the electrons in the bond will spend more time closer to the halogen (e.g. the fluorine, chlorine or bromine atom) The partially positive carbon that results from the polarisation is therefore **electrophilic** and can be attacked by **nucleophiles**. When a suitable nucleophile attacks the carbo-halide bond, it can displace the halogen in a **substitution reaction** to release the halide anion and form a new bond to the carbon, as shown below.





Elimination reactions

An **elimination reaction** is a type of organic reaction in which two substituents are removed from a molecule in either a one- or two-step mechanism. The atoms or groups being removed can be replaced through either substitution or addition reactions.

Activity 1

Name the following compounds (where necessary use the stock notation to indicate the oxidation state of atoms) and calculate their formula masses:

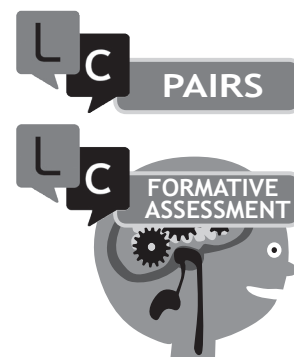
- | | |
|-------------------------------|--------------------------------------|
| a) Fe_2O_3 | b) $\text{Ca}(\text{HCO}_3)_2$ |
| c) SO_3 | d) $\text{K}_2\text{Cr}_2\text{O}_7$ |
| e) LiClO_3 | f) $\text{Mg}(\text{NO}_3)_2$ |
| g) $(\text{NH}_4)\text{SO}_4$ | h) HgO |
| i) AgNO_3 | j) CoBr_2 |
| k) HNO_3 | l) NH_3 |
| m) H_2O_2 | |



Activity 2

Write the formulae for the following compounds:

- | | |
|---------------------------|------------------------|
| a) sulphur (IV) oxide | b) potassium carbonate |
| c) iron (III)nitrate | d) ammonium sulphate |
| e) potassium permanganate | f) aluminium sulphate |
| g) magnesium phosphate | h) carbon (II) oxide |
| i) calcium hydroxide | j) beryllium sulphite |
| k) copper (II) sulphate | l) Ammonium phosphate |





INDIVIDUAL



FORMATIVE ASSESSMENT

Activity 3

Balance the following equations and name the products:

- Which reactions probably represent decomposition by heating?
- State which reactions represent acid-base reactions or redox reactions.
- In the case of redox reactions, state what is being reduced and what is being oxidised.

	Reactants	Product(s)
a)	$K + O_2$	K_2O
b)	H_2O_2	$H_2O + O_2$
c)	$Cl_2 + NaBr$	$NaCl + Br_2$
d)	$SO_2 + O_2$	SO_3
e)	$KClO_3$	$KCl + O_2$
f)	$Cu(NO_3)_2 + C$	$CuO + NO_2 + CO_2$
g)	$Fe + HCl$	$FeCl_2 + H_2$
h)	$Mg + N_2$	Mg_3N_2
i)	$MnO_2 + HCl$	$MnCl_2 + H_2O + Cl_2$
j)	$C_{12}H_{22}O_{11} + O_2$	$CO_2 + H_2O$
k)	$AgNO_3 + BaCl_2$	$AgCl + Ba(NO_3)_2$
l)	$N_2 + H_2$	NH_3 (<i>the Haber process</i>)
m)	$NaOH + H_2SO_4$	$Na_2SO_4 + H_2O$
n)	$Al + O_2$	Al_2O_3
o)	KNO_3	$KNO_2 + O_2$



INDIVIDUAL

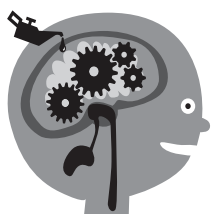


SUMMATIVE ASSESSMENT

Activity 4

Write balanced chemical equations for the following reactions:

	Reactants	Product(s)
a)	magnesium + oxygen	magnesium oxide
b)	copper(II)oxide + carbon	copper + carbon(IV) oxide
c)	iron + hydrogen chloride	iron(III) chloride + hydrogen
d)	sodium + water	sodium hydroxide + hydrogen
e)	copper(II) carbonate	copper(II) oxide + carbon dioxide
f)	chlorine + potassium bromide	potassium chloride + bromine
g)	aluminium + iodine	aluminium iodide
h)	calcium carbonate + an acid	calcium chloride + water + carbon dioxide
i)	sodium hydroxide + an acid	sodium sulphate + water
j)	ammonia + hydrogen chloride	ammonium chloride



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