Chemical reactions

Introduction

This topic explores the key concepts of chemical reactions as they relate to:

- physical and chemical change
- hazardous chemicals
- separation of mixtures
- acids and bases
- oxidation and reduction.

Key concepts of chemical reactions

The activities in this topic are designed to explore the following key concepts:

- A substance that undergoes a physical change remains that substance even though it may have changed state or shape.
- A substance that undergoes a chemical change is no longer the original substance.
- In all chemical reactions, the total mass of the reactants equals the total mass of the products.
- In all chemical reactions, the total number and types of atoms in the reactants equal the total number and types of atoms in the products.
- Food atoms do not disappear in digestion and they do not convert to energy in the body.
- In any chemical or physical change, the total amount of energy remains constant.
- Chemical reactions will not occur until the reactants receive enough energy (activation energy) to break chemical bonds.
- In an exothermic reaction, the total energy of the products is less than that of the reactants—heat is released to the surrounding substances.
- In an endothermic reaction, the total energy of the products is more than that of the reactants—heat is taken from the surrounding substances.
- Chemical reactions are equilibrium reactions; that is, forward and reverse reactions take place at the same time.
- Chemical equilibrium can be disturbed by adding or removing any of the products or reactants, by changing the volume, or by changing the temperature.

- Any change that affects the position of an equilibrium state will cause that equilibrium state to shift, if possible, in such a way as to partially oppose the effect of that change.
- A catalyst is a substance that can alter the rate of a chemical reaction without being used up in the reaction.
- For something to combust or burn, fuel, oxygen, and activation energy input are needed.
- An acid is a proton donor; a base is a proton acceptor.
- Redox reactions involve the transfer of electrons from one reactant to another.

Students’ alternative conceptions of chemical reactions

Research into students’ ideas about this topic has identified the following non-scientific conceptions:

- A chemical is a substance made by humans.
- Mixing an acid with a base (without regard to quantities) neutralises the base, resulting in a neutral solution.
- In neutralisation, all the H^+ and OH^- ions are cancelled.
- Oxidation is the addition of oxygen in a reaction; reduction is the removal of oxygen.
- If a reaction includes oxygen, it is an oxidation reaction.
- Oxidation and reduction reactions can occur independently.
- If water appears during burning, it was present in the wood.
- Smoke formed during combustion was already in the wood.
- Combustion is a change of state of matter: solid or liquid to gas.
- Oxygen aids, but does not participate in, combustion.
- A candle burning is endothermic, because heat is needed to initiate the reaction.
- Chemical reactions are reactions that produce irreversible change.
- Chemical reactions are caused by mixing substances.
- Chemical reactions must be driven by external intervention—for example, heat.
- Energy is used up in a chemical reaction.
- Chemical bonds store energy.
- Energy is created in chemical reactions.
- The products of chemical reactions need not have the same mass as the reactants.
- Water disappears as it evaporates.

Physical and chemical change

There are two basic types of change: physical and chemical. We will now explore the general characteristics of physical and chemical change, which will assist you in understanding the complex changes that are discussed later in this topic.

Characteristics of physical and chemical change

If a substance undergoes a physical change, it remains that substance even though it may have changed state or shape. For example, if you cut up some paper into tiny bits, this represents a physical change; the tiny pieces are still paper, even though cutting the paper means that some of the intermolecular bonds connecting the paper molecules are broken.

All changes of state represent a physical change. Even though water is called something different in each of its three states (ice, water and steam or water vapour), the substance is still water. The water molecules are still intact. In a change of state, the weak intermolecular bonds (also called ‘hydrogen bonds’) between the molecules break (by melting, boiling or sublimation) or re-form (by condensation, freezing or deposition).

Another significant feature of physical change is that most change is reversible; for example, steam can be converted back into liquid water. In the example of the cut paper, it may be difficult to imagine how the bonds can be re-formed, but the paper can be pulped and recycled as paper sheets.

Another example of physical change is a substance dissolving. When you dissolve salt crystals in water, the ionic bonds between the ions break and so the ions are free to move around the solution. One can evaporate the water and the crystals of salt will re-form, thus re-establishing the ionic bonds.

If a substance undergoes a chemical change, a new substance (or substances) is formed. In many cases, the change is irreversible; for example, burning a substance involves the process of combustion, in which oxygen in the air combines with a whole substance or part of a substance to form new substances.

Fossil fuels we use every day undergo combustion when heated. When we burn methane and ethane in natural gas at the jet of a gas stove, propane and butane in a gas barbecue or octane in a car engine, the main products are water and carbon dioxide. Let’s take the specific example of the combustion of ethane. The chemical reaction can be described by the following chemical equation:

$$\text{ethane } (C_2H_6) + \text{oxygen } (O_2) \rightarrow \text{carbon dioxide } (CO_2) + \text{water } (H_2O)$$

In this chemical reaction, the covalent bonds in the ethane and oxygen molecules are broken so that the constituent atoms re-form in other arrangements to produce different molecules: water and carbon dioxide. This reaction is irreversible, which means that carbon dioxide and water when mixed and heated or cooled will not break bonds and re-form as ethane and oxygen. In the combustion of fossil fuels, other chemicals are produced (this is not reflected in the chemical equation given above); for example, an incomplete combustion process often occurs in car engines, where carbon monoxide (CO) and other gases are produced.

Another example of a chemical change is the digestion of food. Following the physical change in the food that occurs when we chew it, chemical reactions start almost immediately when saliva is added to the food. Chemical compounds called ‘enzymes’ help to break down the food into simpler molecules that can be carried in the bloodstream and that dissolve in the water within the cells of the body (i.e. chemical reactions occur that break large molecules into smaller molecules). Enzymes in saliva start the breakdown process and other enzymes and chemicals break down the food in the stomach, duodenum and upper colon. The smaller particles of food enter the bloodstream and are carried to cells in the body, where further chemical reactions occur. Waste products from these chemical reactions are carried by the blood to be expelled by the body.

Conservation of mass in physical and chemical changes

Because we can reverse a physical change, the amount of mass present before the physical change remains the same following the change. In other words, the number of atoms present before the change remains the same following the change. All that occurs in the physical change is the breaking or forming of chemical bonds. This also applies to chemical changes.

In a chemical reaction, chemical bonds between atoms are broken and re-formed in other arrangements; the atoms remain intact. That is, the number and types of atoms before the chemical reaction are the same following the reaction. There is no loss of mass in a chemical reaction. This may be surprising when considering, for example, burning a log of wood. Quite obviously, the weight of the log far exceeds that of the ash that remains after burning; however, if you consider the gases that are produced by the combustion of the log, then the total mass remains the same.

Another interesting effect occurs when steel wool is burnt. The iron (Fe) reacts with oxygen to produce iron III oxide (Fe_2O_3), which is a dark powder, after combustion. The mass of the powder is greater than that of the original steel wool. Can you explain why?

If we define the substances present before a chemical reaction as ‘reactants’ and the substances formed in the chemical reaction as ‘products’, then we can state the law of conservation of mass or atoms as follows:

> In any chemical reaction, the total mass of the reactants equals the total mass of the products. In addition, the number and types of atoms in the reactants equal the number and types of atoms in the products.

Consider the chemical equation for the combustion of ethane:

Reactants		Products
ethane (C_2H_6) + oxygen (O_2)	→	carbon dioxide (CO_2) + water (H_2O)

The equation appears to be at odds with the law of conservation of mass or atoms. The reactants are two carbon atoms (C), six hydrogen atoms (H) and two oxygen atoms, whereas the products are one carbon atom (C), two hydrogen atoms (H) and three oxygen atoms (O). In this situation, one carbon atom and four hydrogen atoms seem to have disappeared, whereas one oxygen atom has appeared.

To maintain the law of conservation of mass or atoms, you need to consider that more than two molecules (one ethane and one oxygen molecule) interact in the first place: multiple molecules are interacting. Consider the following chemical equation:

$$2(C_2H_6) + 7(O_2) \rightarrow 4(CO_2) + 6(H_2O)$$

The large numbers in front of the molecular formulas for each molecule indicate how many molecules are involved in the reaction. Therefore, two ethane and seven oxygen molecules interact to produce four carbon dioxide and six water molecules. For the reactants, there are now four carbon atoms, twelve hydrogen atoms and fourteen oxygen atoms. This is also true for the products. For this chemical reaction, note that the law of conservation of mass or atoms has been obeyed even though new substances have been formed.

A major misconception is that high-energy foods, such as carbohydrates, are literally used up by the body—that by giving us energy these foods disappear in the processes within our bodies. If we are to believe the law of conservation of mass or atoms, this cannot be true. Foods such as carbohydrates undergo chemical reactions within the body, which enhance the delivery and expulsion of molecules or atoms from the cells within it. In a balanced diet, the number of atoms that enter the body equals the number of atoms that leave the body (through respiration, perspiration and excretion). If this doesn’t occur, the person loses weight (more atoms leave than enter) or gains weight (more atoms enter than leave). There is no other way to lose or gain weight. Note that no food atoms disappear in the digestion process—it is just that some foods make the expulsion of atoms from the body more efficient than others (such as excessive amounts of fat, which the body stores). Exercise merely speeds up the process of expulsion of atoms from the body. For example, exercise results in a greater rate of breathing and perspiration.

Energy in physical and chemical changes

Energy is quite an abstract concept and is interpreted in many ways. This may be because the concept is used in many contexts in science and everyday life: energy is part of understanding how electrical circuits work; food packaging lists information on energy content; we talk on an everyday level about, for example, the abundance of energy children have or the lack of energy we have. Another reason that the concept of energy is difficult to understand is that it is not a physical entity. However, you can see the effects of an object that has energy.

For our purposes, energy can be understood in terms of the idea that matter with energy can cause change or movement in other matter. Energy comes in different forms, for example:

- Kinetic energy: A moving car has kinetic energy, which enables it to tow a trailer. Any object that is moving has kinetic energy.
- Heat energy: A hot cup of tea has heat energy, which has the ability to raise the temperature of the cold milk that is added. Heat energy is transferred from hot objects to cooler objects.
- Potential energy: A stretched bowstring has potential energy, which enables it to fire an arrow—that is, it gives the arrow kinetic energy.
- Chemical energy: Petrol has chemical energy, which pushes the cylinders in a car engine when the petrol is ignited. Food has chemical energy, which is converted into kinetic energy in the motion of our blood and limbs. The energy content of food is often provided as information on food packaging, particularly on breakfast cereals.
- Electrical energy: Electric charges flowing in a wire near a magnet have the ability to vibrate the wire, as happens with a loudspeaker.
- Light energy: Considered pure energy, light energy has the ability to raise the temperature of your skin while you lie in the sun. Your body absorbs the light energy that falls on it. This type of energy is sometimes called ‘solar energy’, ‘radiant heat’ or one form of ‘electromagnetic radiation’.

Enthalpy

The total amount of energy stored in a substance is called the ‘enthalpy’ or heat content of that substance. Enthalpy consists of two basic types of energy: kinetic energy and potential energy.

Kinetic energy may be considered to be motion energy; an object in motion has kinetic energy. A particle that has mass and moves has kinetic energy. The kinetic energy of an atom relates to the motion of the electrons and the atoms themselves. The temperature of a substance is a measure of the average kinetic energy of all its component atoms. When a substance rises in temperature, its atoms move faster, and, when it cools, its atoms slow down.

The potential energy of an atom is related to the attractions and repulsions present between protons and electrons within the atoms as well as the attractions present between different atoms. In the section ‘Bonding’ in the topic ‘Chemical substances’, it was established that chemical bonds arise through the overall electrostatic attractions between atoms. The potential energy of an atom relates to the extent of the bonds it forms with other atoms. In bonds, electric charges are attracted to each other: the further they are from each other, the greater their potential energies become. As electric charges are closest for strong bonds, atoms have low potential energy when strongly attached to other atoms and high potential energy when weakly attached to other atoms.

You may think that having high potential energy means there is a strong bond, but this is not true. If we draw a comparison with gravitational potential energy, a brick that is 1 m off the ground has less potential energy than a brick that is 10 m off the ground. If we let each brick fall, the higher falling brick will receive a greater transfer of energy. Similarly, a bowstring has more elastic potential energy when stretched a greater distance. These two examples are consistent with the potential energy between atoms. As the energy involves electric charges, it is sometimes referred to as ‘electric potential energy’. As there are different types of chemical bonding, the potential energy of atoms can vary.

In a physical or chemical change the enthalpy of a substance changes. This means that the kinetic and/or potential energies of the atoms increase or decrease. If the kinetic energies increase or decrease, the substance changes temperature and if the potential energies of the atoms change, the strengths of the chemical bonds change. However, there is a rule that energy is neither created nor destroyed, and so if the enthalpy of one substance increases, there is a corresponding decrease in the enthalpy of other substances involved in the physical or chemical change.

Conservation of energy and energy transfer

Energy is neither created nor destroyed, but can be transferred from one object to another or transformed from one type of energy to another. This means that kinetic energy can transfer from one atom to another, or the potential energy of an atom can be transformed into kinetic energy. Alternatively, the kinetic energy of an atom can be transformed into potential energy, and, vice versa, potential energy can be converted into kinetic energy.

In any physical or chemical change, the total energy (kinetic energy plus potential energy) of all the substances involved in the change remains constant. Energy can transfer or be transformed from one substance to another, but it cannot be created or destroyed.

A substance can increase or decrease its enthalpy (kinetic energy plus potential energy) in two ways:

- through direct contact with substances that contain atoms with more or less kinetic energy than the atoms within it
- by absorbing pure energy, called ‘electromagnetic energy’, which is radiated through space by a distant source.

Transfer of energy by direct contact between substances

The transfer of energy by direct contact can occur in several ways, including bringing a substance into direct contact with a hotter or colder substance, or applying a direct force (for example, contact forces, such as hammering, or frictional forces, such as sliding on a rough surface).

Take the example of a hot nail placed in a bucket of cool water. In this scenario, energy is transferred from the nail to the water. The evidence of this is the increase in the water temperature and the decrease in the nail temperature. On a microscopic level (using the particle model), the nail atoms have greater kinetic energies than the water molecules because of their greater speed and mass (kinetic energy is related to both speed and mass). The nail atoms collide with the water molecules, thus slowing down (losing kinetic energy). This will result in the water molecules speeding up (gaining kinetic energy). As the average kinetic energy of the atoms or molecules in a substance directly relates to its temperature, the nail cools down while the water heats up.

The collisions continue to occur until thermal equilibrium is reached. This means that all the atoms or molecules have the same average kinetic energies. This model of energy transfer explains why a substance cools or heats from the outside in. For example, food heats up from the outside when placed in a conventional oven (microwave ovens use a different energy transfer, which is explained later in this topic).

The kinetic energy of atoms or molecules might not be transferred entirely to the kinetic energy of other atoms or molecules. Some of the energy might be transferred to increasing the potential energy of other atoms or molecules. This will mean that the bonds are stretched or broken; however, the total energy remains constant. The energy lost by one atom or molecule is gained by another atom or molecule.

As you heat up a substance, the transfer of energy increases both the kinetic energies and potential energies of the atoms or molecules. The temperature increases and the bonds begin to stretch. However, at a change of state (melting, boiling, sublimation or evaporation), any further energy transfer goes into breaking bonds. We can imagine the bond breaking when the separation of the atoms or molecules is such that their potential energies reach a critical value and the atoms or molecules become free from each other. This type of energy transfer explains why, when ice melts, the temperature of the ice–water mixture stays at 0 ˚C until all the ice melts. Similarly, the temperature of a water–steam mixture remains at 100 ˚C until all the liquid water boils.

Energy is required to break bonds during a change of state. This explains why you cool down when you perspire, as perspiration evaporates into the air. During evaporation, energy is transferred from your skin (it cools) to change the state of the perspiration (the bonds break). In the evaporation of the perspiration, the energy required to break the bonds comes from the skin. The kinetic energies of the skin molecules decrease, resulting in the skin cooling.

As a substance cools, the transfer of energy is such that energy leaves the substance for surrounding substances. During a change of state (e.g. condensation, freezing, deposition), energy is released as bonds form. The atoms or molecules now have decreased potential energy. The energy transfer to neighbouring substances results in heating. This can explain why steam burns are more serious than boiling-water burns, even though the steam and water are at the same temperature. For steam, energy is transferred when the water changes state (to steam) and is transferred further when the steam cools down. For boiling water, energy is transferred only when it cools down.

For a substance to undergo a chemical reaction, it first must have energy added to break its bonds. This energy is called ‘activation energy’. The activation energy allows bonds to be broken so that the atoms can rearrange themselves to form new substances. For example, a mix of octane (petrol) and oxygen requires a spark to initiate the combustion process that occurs in car engines.

In summary, the enthalpy (total energy = kinetic energies + potential energies of atoms or molecules) of a substance can change when in direct contact with another substance. When a substance heats up (its kinetic energy increases), its bonds stretch and break (its potential energy increases) and its enthalpy increases. In contrast, when a substance cools (its kinetic energy decreases), bonds form and shrink (its potential energy decreases) and its enthalpy decreases. If the enthalpy of one substance increases, there is a corresponding decrease in the enthalpy of the surrounding substance. In this way, energy is conserved; it is neither created nor destroyed.

Transfer of energy by absorption from a distant source

The second way atoms can increase their kinetic or potential energy is by absorbing electromagnetic radiation. There are various types of electromagnetic radiation, such as solar energy (from the sun), microwaves (from a microwave oven), infrared radiation (from hot objects), X-rays (from medical machinery) and gamma rays (from radioactive material). This type of radiation travels from its source through space at the speed of light and when atoms absorb this energy they increase their kinetic energy and/or potential energy. The atoms in the source are not in contact with the atoms that absorb the radiation, so this process is quite different from the one described earlier in the section ‘Transfer of energy by direct contact between substances’.

An example of the absorption of electromagnetic radiation is when we lie in the sun. The solar energy that is emitted from the sun travels millions of kilometres and is absorbed by the atoms in our skin, and so we heat up. In a microwave oven, the water molecules in the food absorb the microwaves, whether they are on the surface of the food or not. This explains why food cooked in a microwave oven heats from the inside. Gamma radiation and X-rays can also penetrate the skin and be absorbed by different internal molecules in the body. This results in an increase in the kinetic or potential energy of the atoms.

Just as atoms can absorb electromagnetic radiation, they can also emit electromagnetic radiation. Some of the kinetic energy of the atoms is converted into a pure form of energy (electromagnetic radiation), which is emitted and may be absorbed by other atoms. Atoms at room temperature give off electromagnetic radiation called ‘infrared radiation’. This type of radiation is emitted in great amounts from fires. Most of the heating of your body on a cold night around a camp fire is a result of the absorption of infrared radiation.

As well as increasing in temperature, a substance can change state (for example, by evaporation) or undergo a chemical change when it absorbs electromagnetic radiation. The absorbed energy is greater than the activation energy of the atoms or molecules. This means that bonds will be broken and a change of state or a chemical reaction will occur. Examples of chemical reactions initiated and sustained by the absorption of electromagnetic radiation include:

- cooking a cake in a microwave oven
- tanning your skin while lying in the sun
- fading materials through exposure to the sun
- production of mutant cells leading to cancer.

In regard to the last example, cell reproduction occurs within the body on a regular basis. These are chemical reactions. Absorbed energy from UV light, microwaves, X-rays and gamma rays induces chemical reactions on or inside the body, with the possibility of producing fast-growing and replicating mutant body cells. X-rays and gamma rays are more dangerous as they carry more energy. However, some studies suggest that the low-energy electromagnetic radiation from electric powerlines and mobile phones can cause cancer-producing cells with long exposure to the radiation.

To be consistent with the law of conservation of energy, a substance that emits electromagnetic radiation loses energy (decreases in enthalpy).

In summary, the transfer of energy from one substance to another can occur at a distance through the emission and absorption of electromagnetic radiation. The enthalpy of the source of radiation decreases while the enthalpy of the absorber increases.

Chemical reactions: exothermic and endothermic reactions

We have explained that a chemical reaction will not occur until substances (reactants) receive enough energy (activation energy) to break chemical bonds. This allows atoms to redistribute themselves to form new bonds and thus form new substances (products). The activation energy needs to be thought of as a barrier to be overcome. If the bonds between the reactants are strong, greater activation energy is required to initiate a chemical reaction; if the bonds are weak, less activation energy is required.

In the first stage of a chemical reaction, the enthalpy of the reactants increases through some form of energy transfer. Therefore, the total energy of the reactants increases by the amount of the activation energies. At this point, bonds are broken. In the next stage, new bonds are made in the formation of the products. The total energy of the products (i.e. the sum of the enthalpies) may be greater than, or less than, that of the reactants. When the total energy of the products is less than that of the reactants, the chemical reaction is called an ‘exothermic reaction’, and when the total energy of the products is more than that of the reactants, the chemical reaction is called an ‘endothermic reaction’.

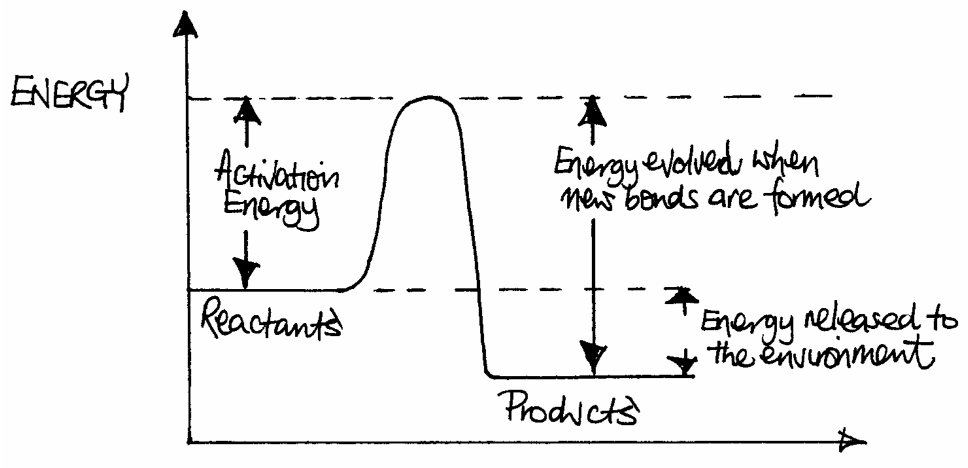
To be consistent with the law of conservation of energy, in an exothermic reaction, excess energy is transferred to the surrounding materials that do not take part in the reaction (the surrounding environment): the environment heats up.

In an endothermic reaction, energy is taken from the surrounding environment: the environment (the surrounding substance) cools down.

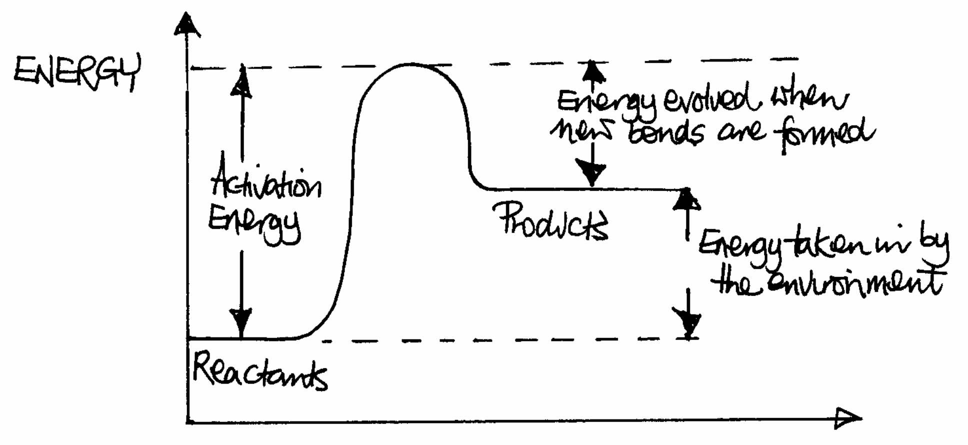
The different types of chemical reaction are shown in Figure 1. Note that the total energy of the whole system (surrounding environment–reactants–products) remains constant before and after the reaction, whereas this is not true for the total energies of the reactants compared to the products.

FIGURE 1: ENERGY DIAGRAMS FOR ENDOTHERMIC AND EXOTHERMIC REACTIONS

Energy diagram for an exothermic reaction



Energy diagram for an endothermic reaction



A good example of an endothermic reaction is the use of an instant icepack. Instant icepacks can be used to treat minor burns as well as sporting injuries, such as sprains. A typical icepack contains the ionic compound ammonium nitrate salt (NH_4NO_3), which reacts with water. In solution (the ionic solid dissolved in water), the ionic bonds are broken, freeing up ammonium ions (NH_4^+) and nitrate ions (NO_3^-). During the reaction, energy is taken from the surrounding environment (for example, the ankle), thus cooling it down. The equation for the reaction is:

$$NH_4NO_3 \xrightarrow{\text{water}} NH_4^+ + NO_3^-$$

Many foods we eat undergo exothermic reactions and literally warm us up. Glucose ($C_6H_{12}O_6$), a type of sugar found in many foods, reacts with oxygen (O_2, in the air we breathe) to produce an exothermic reaction, with carbon dioxide (CO_2) and water (H_2O) as the products. The energy given off assists in the maintenance of a constant internal body temperature, which is important for many processes. The equation for the reaction is:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

It is important to note that, in this reaction, the reactants involve more than two molecules. Every molecule of glucose reacts with six molecules of oxygen. The reaction produces six molecules each of carbon dioxide and water. As an exercise, convince yourself of the law of conservation of atoms by determining that the total number of each type of atom is the same on each side of the chemical equation.

ACTIVITY: ENDOTHERMIC AND EXOTHERMIC REACTIONS

This activity is taken from 'Wondernet activities' <http://www.chemistry.org/portal/a/c/s/1/wondernetdisplay.html?DOC=wondernet%5Cactivities%5Cactivities.html> (viewed 10 March 2005).

The following activity produces an endothermic reaction and an exothermic reaction.

Part A

You will need:

- **3% hydrogen peroxide (do not use a higher percentage)**
- **measuring spoons**
- **3 tspns yeast**
- **a cup (paper or plastic)**
- **a thermometer (use thermometers with red liquid only)**
- **a watch (with a second hand)**
- **at least three people.**

Pour 2 tablespoons of hydrogen peroxide into a cup. Place the thermometer in the cup. Hold the thermometer and the cup so they do not fall over. Read the temperature and record it in a table.

Measure 1 teaspoon of yeast. Have one partner watch the thermometer and another look at the second hand on the watch.

Dump all the yeast into the cup. Gently swirl the cup while one partner calls out the time every 10 s. When each 10 s is called, another partner should call out the temperature. The third partner should record the temperature in the table. What do you observe?

Make a graph of temperature versus time (temperature on the vertical axis and time on the horizontal axis). During what period of time does the temperature change most? When does it change least? Is this an example of an endothermic reaction or an exothermic reaction?

Part B

You will need:

- **vinegar**
- **baking soda**
- **measuring spoons**
- **water**
- **a cup (paper or plastic)**
- **a thermometer (use thermometers with red liquid only).**

Place 2 tablespoons of vinegar in the cup. Put the thermometer in the cup. Hold the thermometer and cup so they do not fall over. Read the temperature and record it in a table.

Measure 1 teaspoon of baking soda. Dump all the baking soda in the cup. Gently swirl the cup while one partner calls out the time every 3 s. When each 3 s is called, another partner should record the temperature in the table. What do you observe?

Graph your results like those in Part A. During what period of time does the temperature change the most? Is this an example of an endothermic reaction or an exothermic reaction?

How do you think adding water to the hydrogen peroxide before adding the yeast would affect how fast the temperature goes up? Try it and see.

Explanatory note: In Part A of this activity, yeast is added to hydrogen peroxide. A chemical in the yeast causes a reaction in which the hydrogen peroxide breaks apart to form oxygen gas and water. This reaction gives off heat. In Part B of this activity, baking soda is added to vinegar and again the temperature changes are recorded. The baking soda reacts with the vinegar to produce carbon dioxide gas and water. This reaction causes a drop in temperature.

Chemical equilibrium: forward and reverse reactions

The descriptions of chemical equations given above assume that when the bonds between atoms are broken new bonds are formed with new arrangements of atoms. Given the explanations, it should be possible for the original bonds to re-form. In other words, the reaction should be reversible. This is what actually happens. However, the degree to which a chemical reaction reverses varies with the nature of the reactants as well as other factors, including the amount of each reactant, the volume of the reactants and the temperature of the environment.

When a chemical reaction is initiated (that is, the activation energy of the reactants is exceeded), the bonds between the atoms break. At this point, two types of reaction can occur, a forward reaction and a reverse reaction. In the forward reaction, new bonds with different arrangements of atoms are produced (the products are different substances from the reactants). In the reverse reaction, the atoms re-form bonds with the same arrangement of atoms as before (the products are the same as the reactants). Initially, the rate of the forward reaction differs from the rate of the reverse reaction. However, after a while, the rate of the forward reaction equals the rate of the reverse reaction. When this occurs chemical equilibrium is established.

Technically speaking, all chemical reactions are equilibrium reactions (i.e. there are forward and reverse reactions) but in many cases the degree of the reverse reaction is so small that it can be effectively ignored. In these circumstances the chemical equations show an arrow in only one direction. However, if the reaction is truly an equilibrium reaction, a double arrow is shown in the equation to indicate that both a forward and a reverse reaction occur. For example, carbon monoxide (CO) and water (H_2O) produce carbon dioxide (CO_2) and hydrogen (H_2). In this reaction, initially CO and H_2O are present, but after a while all four compounds are present.

$$CO + H_2O \leftrightarrow CO_2 + H_2$$

It should be noted that equilibrium reactions never finish. At the point of chemical equilibrium, somewhere in the mixture of reactants and products, reactants are producing products (forward reaction) and, in other places, products are producing reactants (reverse reaction).

Chemical equilibrium can be disturbed (i.e. produce more forward or reverse reactions) in one of three ways:

1 by adding or removing any one of the products or reactants
2 by changing the volume (at constant temperature)
3 by changing the temperature.

Any further discussion on the ways in which equilibrium reactions proceed with these changes can be explored through predictions made from Le Chatelier’s principle, which states that:

> Any change that affects the position of an equilibrium state will cause that equilibrium state to shift, if possible, in such a way as to partially oppose the effect of that change.

Further explanations of this principle and the manner in which it can predict changes are left to you to research. Note, however, that knowing when chemical equilibrium occurs and how much of the reactants and products you have are very important in industries that use chemical reactions to produce substances. It is also important for manufacturers to be able to change the equilibrium to maximise the production of the substances required.

A very important equilibrium reaction occurs in our bodies; this relates to the process by which oxygen is transported from our lungs to cells in our bodies. Although oxygen is soluble in water (blood), this solubility is not high enough to enable the blood to supply adequate amounts of oxygen to our cells. The blood transports the oxygen by a different mechanism: the red blood cells. These cells contain a special protein (a large molecule, its formula is simplified to Hb_4) called ‘haemoglobin’, which reacts with oxygen (O_2) to form oxyhaemoglobin (Hb_4O_8). The equilibrium reaction is written as:

$$Hb_4 + O_2 \leftrightarrow Hb_4O_8$$

In the lungs, the concentration of oxygen is high and, because of this, the forward reaction becomes more significant than the reverse reaction. In this way, the haemoglobin bonds to the oxygen, which is then transported to the tissues and the cells of the body. In these places, the concentration of oxygen is small, so the reverse reaction occurs more prominently. This releases the oxygen from the haemoglobin. The oxygen is then available to react with the chemicals in the cells. The haemoglobin then returns to the lungs to start the cycle again.

Catalysts and chemical reactions

A catalyst is a substance that can alter the rate of a chemical reaction without being used up in the reaction. Catalysts are important in the manufacturing industry as well as in our daily lives. It has been estimated that without catalysts (enzymes) in our bodies, it would take about 50 years to digest a single meal! Enzymes are protein molecules that increase the rate of reactions that occur in living organisms, particularly in breaking up large food molecules into smaller ones.

Although catalysts are usually used to speed up a reaction, they are sometimes used to slow down a reaction. In these situations they are called ‘inhibitors’.

In terms of the mechanisms involved in a chemical reaction, a catalyst lowers the activation energy that needs to be exceeded if a chemical reaction is to occur. Lower activation energy means that the reaction will occur more readily and more often.

Modern cars are required by law to control their waste by using catalytic converters. These are fitted to exhaust systems. The catalysts used, usually the elements platinum or rhodium, speed up the conversion of carbon monoxide (CO) gas to carbon dioxide (CO_2), and nitrogen oxide (NO) gas to nitrogen (N_2). Carbon monoxide gas is produced by the incomplete combustion process in the engine where octane (C_8H_{18}) reacts with oxygen (O_2). Nitrogen oxide comes from the combustion of nitrogen and oxygen when the temperature of the engine reaches 2000 ˚C.

General types of chemical reactions

Bonds are broken and formed in all chemical reactions. There are two other major types of chemical reaction that will be dealt with in this topic. These are acid–base reactions and reduction–oxidation (or ‘redox’) reactions.

Acid–base reactions involve the transfer of protons (in the form of hydrogen ions, H^+) from one of the reactants (acid) to the other (base). Redox reactions involve electron transfer from one of the reactants (reductant) to the other (oxidant). Acid–base and redox reactions will be explored in further depth later in this topic in the sections ‘Acids and bases’ and ‘Oxidation and reduction’, respectively.

Hazardous chemicals

Chemicals for warfare

Chemical scientists have produced many different types of chemicals that have proven to be hazardous to humans and their environment. Military scientists have concocted a number of dangerous chemicals (e.g. mustard gas) for warfare. These chemicals have been used in the past and the threat of their continued use remains today. To find out more about some of the known chemical agents used in warfare and their effects, access the site <http://www.scc.rutgers.edu/njh/WW2/Chemicals2.htm>.

Persistent organic pollutants

While scientists have deliberately made synthetic substances that do harm, they have also produced substances that were initially designed to make life easier for humans. However, we now know that some chemicals that are designed for use as pesticides or are unwanted industrial by-products (i.e. they are waste products of some chemical process used by industries) are quite hazardous to humans and the environment.

In 1997 the United Nations Environment Programme (UNEP) agreed to phase out chemicals called ‘persistent organic pollutants’, generally known as ‘POPs’. Currently, the focus is on twelve of the worst, known as the ‘dirty dozen’.

POPs are synthetically made, so the term ‘organic’ is somewhat misleading. An organic chemical is any substance whose molecular structure is based on carbon and hydrogen atoms. One or more of the hydrogen atoms, however, may be replaced in the chemical structure by other elements such as sulfur, oxygen, nitrogen, phosphorus, chlorine, fluorine or various metals.

POPs are dangerous for the environment because they persist for long periods and are toxic to animals (including humans). POPs enter the food web via organisms. They don’t dissolve readily in water but dissolve easily in fats, so POPs accumulate in the fatty tissue of living organisms (it is often difficult for the organism to break them down or excrete them). Because of this, POPs undergo a process called ‘bioaccumulation’. This means that they build up over time in the tissues of living organisms. Biomagnification also occurs. This is the process by which organisms that have accumulated POPs in their bodies are eaten by other organisms higher up on the food chain. Toxic substances that are present in the prey species become concentrated in the predator species. Humans, at the top of the food chain, are at greatest risk of accumulating POPs.

POPs have been found in all parts of the globe because of their ability to travel long distances. They evaporate slowly, which allows them to enter the air, travel long distances and settle back to earth. The colder the environment, the less they evaporate, so they are now concentrating on the polar regions of Earth. POPs are toxic to humans even at very low concentrations. The following table lists the chemical names of the ‘dirty dozen’. Only brief details about the common uses and health effects are given. Complete the activity *Investigating POPs.*

TABLE 1: CHEMICAL NAMES OF POPS

POP	**Common use**	**Health effects**
Aldrin	Pesticide	Similar to effects of DDT (below)
Dieldrin	Pesticide	Similar to effects of DDT (below) although fifty times more toxic
Dichlorodiphenyl trichloroethane (DDT)	Pesticide	Reproductive failure in wildlife; links to breast cancer, nervous system disruption and muscular weakness; associated with suppression of the immune system
Endrin	Pesticide	Similar to effects of DDT
Heptachlor	Pesticide	Similar to effects of DDT
Chlordane	Pesticide	Probable human carcinogen (cancer-producing substance)
Hexachloro-benzene (HCB)	Pesticide and industrial by-product	Probable human carcinogen
Mirex	Pesticide	Possible carcinogen; associated with suppression of the immune system
Toxaphene	Pesticide and industrial by-product	Possible carcinogen
Polychlorinated biphenyls (PCBs)	Industrial chemical and by-product	Adverse effects on wildlife and acutely exposed human populations; effects on growth of foetuses; long-term effects on intellectual function
Dioxins	Industrial by-product	Developmental, endocrine and immune system toxicants that may have reproductive effects on humans
Furans	Industrial by-product	Similar to effects of dioxins

ACTIVITY: INVESTIGATING POPS

Access the following three sites that relate to POPs (or find others) and consider the questions that follow:

Greenpeace: 'POPs: Poisoning our planet'
http://www.greenpeace.org.au/toxics/pops/index.html

Physicians for Social Responsibility (PSR): 'Toxics and health'
http://www.envirohealthaction.org/toxics/pollution/

United Nations: 'Persistent organic pollutants'
http://www.chem.unep.ch/pops/

- **What evidence is there that POPs travel to all parts of Earth?**
- **How do scientists know that POPs are accumulating at the polar regions?**
- **What industries have POPs as industrial by-products?**
- **Can these industries use different chemicals?**
- **How serious is the problem?**
- **Are any countries (including Australia) still producing POPs?**
- **What is Australia doing about the problem?**
- **What is the rest of the world doing about the problem?**
- **What strategies are there to resolve this problem?**

Hazardous chemicals in everyday life

Communities use hazardous chemicals on a regular basis and for a variety of reasons. Storage of these chemicals needs to be secure and their location needs to be known to other members of the community, particularly emergency services such as the fire brigade, police and state emergency service (SES). Organisations or individuals that use and store hazardous chemicals follow strict regulations about their use, storage, transport and disposal. There are nine classes of hazardous chemicals (regulations call them 'dangerous goods') and each class has a specific identification logo required to be placed on containers storing the chemicals and their location. An example of the class known as 'explosives' is shown in Figure 2.

FIGURE 2: DANGEROUS CHEMICAL LOGO—EXPLOSIVES

The logo designs of the other eight hazardous chemicals classes can be seen at <http://www.hie.hl.com.au/hazchem.htm>.

The nine classes of dangerous goods

Table 2 lists details about each class of hazardous chemical (or dangerous good).

TABLE 2: CLASSES OF HAZARDOUS CHEMICALS

Class	**Class name**	**Examples**
1 Explosive	1 Explosives	Gunpowder, gelignite, fireworks, fuses, detonators
2 Gas	2.1 Flammable gases	Acetylene, LPG
	2.2 Non-flammable, non-toxic gases	Air, argon, liquid oxygen
	2.3 Poison gases	Nitrogen dioxide, chlorine, anhydrous ammonia, methyl bromide
3 Liquid	3 Flammable liquids	Petrol, kerosene, paint thinners
4 Solid	4.1 Flammable solids	Sulfur, nitrocellulose, picric acid
	4.2 Dangerous when wet	Calcium carbide
5 Oxidising or organic	5.1 Oxidising substances	Calcium hypochlorite (swimming pool ‘dry chlorine’), hydrogen peroxide
	5.2 Organic peroxide	Di-benzoyl peroxide, methyl ethyl ketone peroxide
6 Poison	6.1 Toxic substances	Sodium cyanide, arsenic trioxide
	6.1 Harmful substances	Lead acetate dichloromethane
7 Radioactive	7 Radioactive substances	Uranium, plutonium
8 Corrosive	8 Corrosives	Hydrochloric acid, sodium hydroxide
9 Miscellaneous	9 Miscellaneous	These substances do not have to be licensed

Emergency action code

The dangerous goods listed in Table 2 are not hazardous in their normal use. However, they can become quite hazardous in an emergency situation such as a fire, a gas or liquid leak or spillage, or an explosion. In these circumstances it is important for emergency services to know what chemicals are involved and how to resolve the emergency situation. There is a certain code, called the ‘Hazchem Code’, which conveys information to emergency services personnel where prompt action in the absence of full information could control potentially dangerous situations. This code is usually placed on the outside of a building or grounds that contain hazardous chemicals. Transport that carries hazardous chemicals must also display this code.

The Hazchem Code consists of an alphanumeric system of two or three symbols. The first symbol is always a number from 1 to 4 and designates the equipment suitable for firefighting and, where appropriate, for dispersing spillage (see Table 3).

TABLE 3: HAZCHEM CODE: NUMERIC DISPLAY

Code	**Type of extinguisher**
1	Water jets
2	Water fog (if unavailable, fine water spray may be used)
3	Foam
4	Dry agent (for substances where contact with water is hazardous)

The second symbol is always a letter from P to Z (excluding Q, U and V). This symbol denotes whether there is likely to be a violent reaction or explosion, what protective clothing would be required and whether the spill should be contained or diluted.

TABLE 4: HAZCHEM CODE: ALPHA DISPLAY

Code	**Violent reaction**	**Protective clothing or equipment**	**Environment**
P	Yes	Full protective clothing	Dilute
R	No	Full protective clothing	Dilute
S	Yes	Breathing apparatus	Dilute
S	Yes	Breathing apparatus for fire only	Dilute
T	No	Breathing apparatus	Dilute
T	No	Breathing apparatus for fire only	Dilute
W	Yes	Full protective clothing	Contain
X	No	Full protective clothing	Contain
Y	Yes	Breathing apparatus	Contain
Y	Yes	Breathing apparatus for fire only	Contain
Z	No	Breathing apparatus	Contain
Z	No	Breathing apparatus for fire only	Contain

Sometimes there is a third letter, ‘E’, added to the code. The letter ‘E’ is added when evacuation of the people from the neighbourhood of an incident should be considered by the emergency service. Even if there is more than one type of hazardous chemical, each with a different code, a single code is used. This single code is determined on the basis that the fire-extinguisher used will extinguish each of the chemical fires and the protective clothing or environment response is appropriate for each chemical present. (In Table 4, the shaded areas indicate specialisation for fire.)

Fires and fire safety

Combustion is the burning of a material that involves a chemical reaction of a substance with oxygen. More discussion about these types of reactions will be discussed later in the unit.

There are three requirements for something to combust or burn. These are:

- fuel
- oxygen
- activation energy input (sometimes called ‘ignition temperature’).

Without any one of these components, a fire will not be sustained. By knowing this, you can produce and maintain a fire in addition to putting one out. To make a fire, you need an appropriate substance that is flammable—that is, the substance will burn. In addition to being flammable, you do not want to use a fuel that produces toxic products in the combustion reaction, particularly if they are in a gas state. For example, the burning of treated pine logs and some plastics produces toxic substances.

Given that combustion involves a reaction between oxygen molecules and fuel molecules, to initiate and maintain a fire the oxygen molecules need to interact with the fuel molecules. If the fuel is in a solid state then combustion can only occur at the surface of the material. Therefore, to enhance the chance of creating and maintaining a fire, the surface area of the fuel needs to be as large as possible. This explains why it is easier to light a wood fire if you begin with small shavings or kindling. If the fuel is in very small particles, like coal dust or flour dust, combustion may result in an explosion. Grain silos and coalmines have been known to explode because the coal or grain dust has undergone combustion from something as simple as a spark.

The section ‘Physical and chemical change’ in this topic states that a chemical reaction does not occur until a substance receives an amount of energy in excess of its activation energy. For some fuels, the activation energy may be high (wool garments) or it may be low (nylon garments). The transfer of energy may be undertaken in a number of ways—for example, rubbing two sticks together, which creates heat. Rubbing also occurs with the use of matches. The phosphorus at the end of the match has low activation energy, so rubbing the end of the phosphorus ignites it. This reaction is exothermic, which means energy is released to the environment. This energy is greater than the activation energy of the wood in the match, so it burns.

A fire is easy to maintain once it is going if fuel and oxygen are present. As combustion is an exothermic reaction, which transfers energy to the environment, the condition of activation energy is met for the unburnt wood.

To extinguish a fire you need to take away one of the conditions necessary for combustion. The following examples are ways in which fires are extinguished:

- Throwing water on a wood fire lowers the temperature of the wood so that activation energy is not reached. A covering of water also stops oxygen reaching the wood, and water is not itself flammable.
- Placing a lid over a fat fire in a saucepan in the kitchen stops oxygen reaching the combustion site. Throwing water on the fire may spread the fire as fat is insoluble and will float on top of the water. As the water spreads so will the fire.
- Throwing sand or soil on a camp fire stops oxygen reaching the combustion site. Sand or soil is not flammable.

To extinguish an electrical fire, you need to use a material that is not a conductor of electricity. Some fire-extinguishers use carbon dioxide or dry chemicals. In the event of a power pole fire (with live wires above), fire brigades sometimes spray water in a mist form (small droplets) even though the wires above may be live. If a full spray of water is used, this creates an electrical conducting path to the firefighter holding the hose.

When using a portable fire-extinguisher (every home should have at least one) you need to know if the extinguisher is suitable for the type of fire you wish to extinguish. There are five general types of fires and an equal number of types of fire-extinguishers. Fire-extinguishers are recognised by their logo indicators and the colour of their containers. Table 5 (constructed from information at http://www.hanford.gov/fire/safety/extingrs.htm#fetypes, viewed 10 March 2005) lists the types of extinguishers and fires. The table also indicates which types of fire-extinguisher are suitable for which types of fire.

TABLE 5: FIRE TYPES AND SUITABLE FIRE-EXTINGUISHERS

Fire type	**Type of fire-extinguisher**				
	Contents are electrically conductive		**Contents are not electrically conductive**		
	Water (red)	Foam (blue)	Dry chemical (red with white band)	Carbon dioxide (red with black band)	Halon (yellow)
A ordinary combustibles (wood, paper etc.)	Yes	Yes	Yes	Yes	Yes
B flammable liquids	No	Yes (special foam required for alcohol fires)	Yes	Yes	Yes
C flammable gases	No	No	Yes	Yes	Yes
D combustible metals	No	No	No	No	No
E fire involving live electrical equipment	No	No	Yes	Yes	Yes

For more information on fire-extinguishers, refer to the websites <'All you ever wanted to know about fire-extinguishers' <http://www.hanford.gov/fire/safety/extingrs.htm>

The following activities relate to physical and chemical change and can easily be done at home and within the classroom. As you complete an activity consider the chemical concepts involved. Use the models we have discussed thus far to picture in your mind what is happening on a microscopic scale.

ACTIVITY: WHERE DID THE SUGAR GO?

You will need:

- **a clip-lock bag**
- **vinegar**
- **sugar.**

Measure a quarter of a cup of vinegar into the clip-lock bag.

Add a flat teaspoon of sugar and shut the bag. Mix together the sugar and vinegar.

- **What do you observe?**
- **Where has the sugar gone?**
- **Do you suspect a physical or a chemical change has occurred?**
- **Why?**

Explanatory note: The sugar dissolves. This is a physical change and is, in principle, reversible.

ACTIVITY: FROTH AND BUBBLES

You will need:

- **a clip-lock bag**
- **vinegar**
- **baking soda.**

Measure a quarter of a cup of vinegar into the clip-lock bag. Add a flat teaspoon of baking soda and quickly shut the lid.

- **What did you observe?**
- **What happens to the baking soda?**
- **Where has it gone?**
- **What happens to the vinegar?**
- **Has it changed?**
- **How do you know?**
- **Is this activity an example of a physical or a chemical change?**

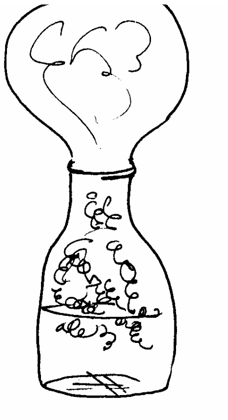
Explanatory note: This is a chemical reaction between the baking soda and the vinegar. It is an acid–base reaction (see the section ‘Acids and bases’ later in this topic), which forms water and a salt (sodium acetate), which remains in solution, and carbon dioxide, which is a gas. The evidence that a chemical reaction has occurred is the presence of the gas and that the mixture no longer smells strongly of vinegar (this is because much of the vinegar has been used up in the reaction).

ACTIVITY: BALLOON BLOW-OUT

You will need:

- **a small soft-drink bottle**
- **a balloon**
- **a spoon**
- **a funnel**
- **vinegar**
- **baking soda.**

Pour vinegar into the soft-drink bottle to a depth of 2–3 cm. Put a teaspoon of baking soda into a balloon (use the funnel) and fit the mouth of the balloon tightly over the mouth of the bottle. Up-end the balloon to allow all the baking soda to fall into the vinegar.



- **What causes the balloon to blow up?**
- **What has happened to the baking soda and the vinegar?**
- **What is in the balloon?**

Take the balloon carefully off the bottle, sealing it and tying it. Blow up another balloon to the identical size. Both balloons are now full of gas. Compare the balloons.

- **Do they feel the same?**
- **Do they drop to the floor at the same rate?**

Explanatory note: Carbon dioxide gas is denser than air and so will fall to the earth. Carbon dioxide is not flammable and so is a good fire retardant as it forms a blanket layer over a combustion site.

ACTIVITY:
CORK ROCKET

You will need:

- **a glass soft-drink bottle**
- **a tight-fitting cork**
- **a spoon**
- **a tissue**
- **vinegar**
- **baking soda.**

Pour vinegar into the soft-drink bottle to a depth of 2 to 3 cm. Carefully pour in a teaspoon of baking soda and put the cork on tightly.

- **What do you think will happen?**
- **Can you explain why the cork flies off?**

Explanatory note: The production of a gas needs more space.

ACTIVITY:
CRAZY
CURRANTS

You will need:

- **a small jar or glass**
- **a spoon or icy-pole stick**
- **vinegar**
- **currants**
- **baking soda.**

Fill the small jar with vinegar, so it is three-quarters full. Put in six currants.
Do they sink or float?

Use the icy-pole stick to put in a small amount of baking soda and stir the mixture. Look carefully at what happens to the currants.

- **Why are the currants going up and down?**
- **What causes the bubbles?**
- **What is in the bubbles?**
- **What has happened to the baking soda? Is it still there?**
- **What has happened to the vinegar?**
- **Once the bubbles stop, how can you make them again?**

Explanatory note: Bubbles of carbon dioxide form and attach to the currants and act as floats that raise the currants to the surface. On reaching the surface, the bubbles burst leaving the currants to sink to the bottom once again. This pattern repeats many times.

ACTIVITY: SALINE

You will need:

- **icing sugar**
- **citric acid crystals**
- **baking soda**
- **a glass**
- **a spoon.**

Make up a powder of 1 teaspoon of icing sugar, half a teaspoon of citric acid crystals and quarter of a teaspoon of baking soda.

Mix these together and stir a teaspoon of the mixture into a glass of water.

- **What happens?**
- **Where has the powder gone?**

Try a small amount of each of the ingredients in some water by itself.

- **What happens in each case?**
- **Which combination of powders causes the fizz?**

Explanatory note: The reactants are citric acid and baking soda, which create carbon dioxide once they dissolve in water and are able to react.

ACTIVITY: SHERBERT RECIPE

You will need:

- **icing sugar**
- **citric acid crystals**
- **baking soda**
- **a glass**
- **an icy-pole stick.**

Make up a powder of 1 teaspoon of icing sugar, half a teaspoon of citric acid crystals and quarter of a teaspoon of baking soda. Mix these together and use the icy-pole stick to put a small amount of the mixture on your tongue.

- **What do you think causes the fizz?**
- **Do any of the powders on their own cause a fizz?**

Explanatory note: This is the same reaction as in the activity *Saline*.

ACTIVITY: ALKA-SELTZER ACTION

You will need:

- **a small jar with a candle that comes to 1 cm below the rim**
- **an Alka-seltzer tablet**
- **a candle.**

Place the candle in the jar, which has water at the bottom. Light the candle. Break the tablet into four pieces and carefully drop the pieces into the water.

- **What do you see?**
- **What causes the bubbles?**
- **What is in the bubbles?**
- **Why does the candle go out?**

Explanatory note: Carbon dioxide is produced once again. The ingredients include citric acid and sodium bicarbonate that react when dissolved in water. The carbon dioxide builds up (it is denser than air) until it covers the burning wick. The flame extinguishes because carbon dioxide, which is non-flammable, forms a blanket between the wick and the oxygen.

ACTIVITY: NAKED EGG

Key idea: The shell of an egg can be removed without breaking it.

You will need:

- **a jar with a lid**
- **a raw egg**
- **a bottle of clear vinegar.**

Place the egg in the jar and fill with vinegar. Close the lid of the jar. Observe what happens over a period of 24 h.

Explanatory note: Bubbles will form over the egg immediately and will increase with time. After 24 h the shell will be gone, and portions of it may be seen floating on the surface of the vinegar. The egg remains intact as the thin membrane stays. The yolk can be seen through the membrane. The change is caused by the acetic acid (CH_3COOH) in the vinegar, which reacts with the eggshells, which are made of calcium carbonate ($CaCO_3$). Two of the products of the reaction are carbon dioxide and water.

ACTIVITY: BREAKDOWN

Key idea: The following experiment breaks up hydrogen peroxide into water and oxygen with the aid of a potato.

You will need:

- **hydrogen peroxide**
- **a raw potato**
- **a glass.**

Fill half the glass with hydrogen peroxide. Add a slice of raw potato and observe what happens. Look particularly for evidence of gas being given off.

Explanatory note: Raw potatoes contain the enzyme catalese. Enzymes are chemicals (also called 'catalysts') found in living cells that speed up the breakdown of food particles. The catalese from the potato causes the hydrogen peroxide (H_2O_2) to quickly break apart into water (H_2O) and oxygen gas. A good test for the presence of oxygen is to bring a flame (from a match) nearby. The flame will flare up. Try it out.

ACTIVITY: GREEN BLOBS FROM STEEL WOOL

Teaching note: In this experiment you will make a green, jelly-like blob by mixing two liquids. One liquid is made by dissolving steel wool. This activity is from van Cleave 1989, p. 102.

You will need:

- **vinegar**
- **'pure' steel wool (i.e. without soap)**
- **household ammonia**
- **two small baby-food jars.**

Fill one jar halfway with steel wool. Add enough vinegar to cover the steel wool, and close the jar. Write 'iron acetate' on the side of the glass. You will have to be patient and let this jar stand undisturbed for 5 days. Then pour 1 tablespoon of your liquid iron acetate into the second jar.

Add 1 tablespoon of household ammonia and stir. What happens? A dark-green, glutinous material forms immediately.

The iron in the steel wool combines with the vinegar (weak acetic acid) to produce iron acetate. The active component of household ammonia is a compound called 'ammonium hydroxide'. A chemical reaction occurs as soon as these two liquids combine. So, ammonium hydroxide + iron acetate makes ammonium acetate + iron hydroxide. Note that what happens is really just an exchange of materials. The same ingredients—ammonium, iron, hydroxide and acetate—are present, but their recombination produces a totally different product. In fact, two liquids (solutions of iron acetate and ammonium hydroxide) have reacted to form a new state—your gel!

The safety concerns in this activity are typical: no eating; wash any skin contact with water; and dispose of materials down the sink with water.

Separating mixtures

A mixture contains two or more substances that are in contact with each other. There are six general types of mixtures:

- solid–solid (concrete)
- solid–liquid (muddy water)
- solid–gas (smoke)
- liquid–liquid (ink)
- liquid–gas (soft drink)
- gas–gas (air).

This section deals with techniques to separate substances contained within mixtures.

Techniques for separating mixtures

Filtration

Filtration occurs when solutions pass through a filter. It is often used in conjunction with sedimentation and decantation. Sedimentation occurs when solids fall and settle on the bottom of solutions. Following sedimentation, decantation occurs when the liquid part of the solution is poured off. A juicer is an example of filtration in the home. The filter (sieve) at the top of the juicer lets the orange juice pass through leaving the pips and pulp behind.

Crystallisation

Crystallisation occurs when a solid in a solution grows into well-formed shapes, called ‘crystals’, mainly by evaporating or boiling off the liquid. Crystallisation was discussed in the topic ‘Water’. The production of refined sugar is an example of crystallisation.

Centrifugation

Centrifugation separates mixtures through a spinning motion in a machine called a ‘centrifuge’. As the centrifuge spins, the heavier parts of the mixture rotate at a greater radius. Sometimes there are holes in the centrifuge that allow the liquid part of the mixture to separate. Alternatively, decanting the liquid from the solid completes the separation. Swagmen used centrifugation when they swung a billy can in a vertical circle so that the tea leaves would separate from the liquid tea.

Adsorption

Adsorption occurs when one component of a mixture (gas or liquid) is attracted to a material that the mixture passes through and forms a layer on the surface of that material. An example of adsorption is when air is passed through charcoal in a gasmask. The dangerous gases collect on the surfaces of the charcoal particles and the clean air passes through.

Chromatography

Chromatography is a technique in which a mixture is placed into or onto a medium and a solvent is passed through the treated medium. As the solvent passes through, the substances in the mixture are carried along. The lighter substances get carried longer distances than the heavier ones, causing them to separate. In forensic work, evidence from a crime scene is analysed for its components using chromatography.

Distillation

Distillation is the process of separating one substance from another when the substances have different boiling points. When a mixture of substances is heated, the substance with the lowest boiling point will evaporate first, leaving the other substance behind. The evaporated substance can be retrieved by passing it through a tube called a ‘condenser’, which is cooled by running water. An example of distillation occurs in the oil industry: the crude oil that is pumped from the ground is a mixture of gas, kerosene, petrol, lubricating oils etc.

Flotation

Flotation is a process in which a substance is added to a mixture that attaches itself to one of the substances in the mixture. This new substance attracts bubbles of air. As the air bubbles rise to the surface, they bring one of the substances in the mixture with them. The bubbles with the attached substance can then be scraped off. Flotation is used in the mining industry to separate metallic ores from rock.

Magnetism

Magnets are substances that attract certain magnetic substances such as iron and steel. Therefore, placing a magnet in a mixture that contains a magnetic substance will result in it being attracted to the magnet, leaving the non-magnetic material behind. Magnetism is used in recycling places to separate iron from other rubbish.

Electrostatic separation

Electrostatic separation occurs when a solid is mixed with a gas and the solid particles are given an electric charge. By then applying an opposite electric charge to the mixture, you can attract the charged solid particles, leaving the gas behind. Electrostatic separation occurs in industrial chimneys to separate the dust and soot from the clean air.

Gravity separation

Gravity separation occurs when one of the substances in a mixture of solids has a higher density than the other. If you add a liquid to the mixture and stir it, the denser solid will gravitate to the bottom first. The less dense solids can then be poured or scraped off leaving the denser solid behind. Gravity separation occurs when you pan for gold. Water is added to crushed rock in a pan and the mixture is swirled about in the pan allowing any gold pieces (which are denser than rock) to collect at the bottom of the pan.

ACTIVITY: SEPARATING MIXTURES

Teaching note: You can explore the methods for separating mixtures in a variety of ways, from giving students specific directions about how to separate mixtures to providing students with an unknown mixture to separate as a problem-solving task.

Consider the following mixtures (the equipment required is not given; students will need to determine this themselves):

- **soil from the garden (a mixture of different soils and organic matter such as twigs and leaves)**
- **salt and chalk**
- **moist sand (a mixture of water and sand)**
- **salt and pepper**
- **immiscible liquids, such as glycerine, brine (salty water) or sugar syrup, water, alcohol (e.g. methylated spirits), or vegetable oil; any combination of these will do; the effect is more striking if each liquid is coloured with food colouring first**
- **iron filings, sand and salt**
- **salty water.**

There are multiple ways to separate the mixtures above. Here are some examples that use everyday household items:

Soil: Pick out the large pieces of twigs, leaves or larger pebbles by hand. Place the soil in a glass jar and add water. Shake and/or stir the mixture and then let it stand for some hours. Gravity separation will occur. The liquid can be decanted.

Salt and chalk: Add water to the mixture. The salt will dissolve (it is soluble) but the chalk will not (it is insoluble). You can then filter the mixture (coffee filters placed in a plastic funnel work fine). Place the salty water in a tray and leave it on a windowsill. Over the course of a sunny day the water will evaporate.

Moist sand: Put the sand in a bucket and then place an open jar in the middle of the sand. Place some clear plastic loosely over the top of the bucket. Put a stone or small heavy object in the centre of the plastic so that it is taut but in a conical shape (the plastic should not touch the sand). You will need to seal the plastic so that the bucket is airtight. Place the bucket in direct sunlight and distillation will occur. The evaporated water condenses on the plastic. It then falls towards the centre and drops into the jar.

Salt and pepper: Rub a plastic ruler with some cloth and bring it over the mixture. Rubbing the ruler makes it electrically charged. The pepper will be electrically attracted to the charged ruler through a phenomenon called ‘electrostatic induction’. If any pepper remains, you can follow the process described in ‘salt and chalk’ above as salt is soluble in water and pepper is not.

Immiscible liquids: All the liquids have different densities and are insoluble in each other (‘immiscible’) so will form layers when placed in a container. The denser layer will be on the bottom. To separate the mixture, place it in a plastic soft-drink bottle. Place a hole in the bottom of the bottle and cover it. If the hole is small enough and you are careful, by opening and closing the hole the various layers can be poured into different containers.

Iron filings, salt and sand: First use magnetism to extract the iron filings from the mixture. The residual mixture can be separated by the method described in ‘salt and chalk’ above.

Salty water: If you do not want to retrieve the water, simply place the mixture in a flat pan and leave it on a windowsill for a few days. Otherwise, you could try a similar technique to that for ‘moist sand’ above.

ACTIVITY: CHROMATOGRAPHY

You will need:

- **filter paper**
- **plastic (transparent) cups or glasses**
- **saucers**
- **water.**

This chromatography technique will separate the pigments in colours used in inks and coverings for lollies. You can use any of the following: water-soluble coloured textas, M&M’s, Smarties, jelly beans. A number of approaches are listed below. Try some out to determine which give the best result.

Smarties or jelly beans

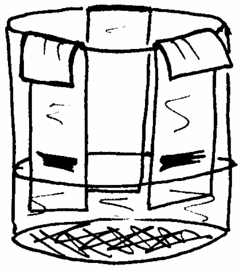
Scrape the coloured coating off a number of lollies of the same colour and discard the rest of the lolly. Crush the coating (use a mortar and pestle), adding a little water to make a coloured liquid. Place a drop of the liquid into the centre of a piece of filter paper. Add a drop of water to the previous drop. The colours should separate on the filter paper, making different rings according to the different food colourings used in the lollies.

Coloured water-soluble textas

Try black and green textas first. There are two methods of setting up your chromatography paper.

Method 1:

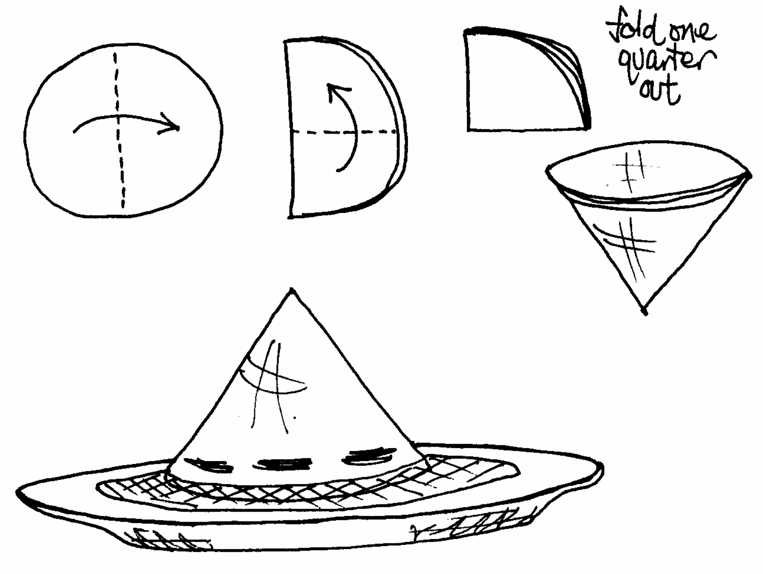
Cut strips of filter paper approximately 10 cm long and 3 cm wide (write the name of each colour you use near the end of each strip). Pour approximately 2–3 cm of water into a glass or transparent plastic cup. Approximately 2 cm from the end of each strip make a mark with a coloured texta. Carefully place your strips in the cup so that only a small portion of the bottom of each strip touches the surface of the water. Be sure that the coloured mark from the texta is above the water surface (see the figure below). Allow the paper to stand for 1 h.



Method 2:

Using circular filter paper, fold the filter paper in half and then in half again. Draw texta marks approximately 2 cm from the curved edge (see the figure below). Make sure the marks do not touch each other. Fold out one side of the filter paper to make

a cone and secure with a paperclip. Fill a saucer with water and place the rounded edge of the filter paper in the water. Allow the paper to stand.



Explanatory note: It takes approximately 1 h for the colours to separate. A trail of blue, yellow and purple is seen from the black mark, and the green mark produces a trail of blue and yellow. Try out textas of other colours as well.

Smarties or M&M's

Soak the Smarties or M&M's in a little water. Use method 1 or 2 described above to set up your chromatography test. To make the coloured mark on the filter paper place a dot of coloured liquid on the filter paper (2 cm from end or edge), dry it with a hair dryer and repeat several times. Another way to get the colour onto the paper is to lick the edge of the Smartie or M&M and draw a line with it on the filter paper. Repeat this on the same spot on the filter paper to build up a dense colour.

Non-water-soluble inks

The colours from permanent pens do not separate in water, as the ink does not dissolve in it. Therefore, instead of water, use other solvents, such as rubbing alcohol or methylated spirits. Alcohol is flammable and easily forms vapours, so conduct your tests in well-ventilated areas.

Crime lab chemistry

The following is detailed in Barber 1989b. Chromatography is used in crime labs to separate the components of 'clue' substances, such as blood, ink and other mixtures found at crime scenes. Chromatograms of these clue substances (strips of filter paper with the separated mixtures) are then compared with those of suspected sources. The students are to determine which of several different pens wrote a ransom note. The pens need to be of the same colour (red or black work well) but different brands. Some can be oil-based and others water-based.

Companies use different combinations of pigments to produce their coloured inks and so different pens will produce different chromatograms. Students will need to produce a chromatogram from the ink on the ransom note; it is best if the ransom note is on filter paper (paper towelling might also work) and a strip can be cut from it (see method 1 in the activity *Chromatography*). Students then produce chromatograms for the different pens found. They will soon work out if water or alcohol is required as a solvent.

ACTIVITY: PLANT PIGMENTS

You can use chromatography to test plant pigments. Transfer the plant pigment to the filter paper by placing the edge of a coin over the sample just hard enough to crush the plant. Repeat several times to build up a dense colour. Try spinach, brightly coloured flower petals, purple cabbage and other plants. Investigate which solvents cause plant pigments to move up the paper (vinegar, water, methylated spirits and alcohol).

ACTIVITY: CRYSTALS

Teaching note: This technique, adapted from van Cleave 1989, pp. 138, 142, will produce crystals of Epsom salts and salt.

Key idea: The following techniques can be used to make a variety of different crystals.

You will need:

- **two saucers**
- **dark cardboard**
- **Epsom salts**
- **salt**
- **two small baby-food jars with lids**
- **a tablespoon**
- **scissors.**

Half fill each jar with water and add 2 tablespoons of Epsom salts to one jar and 1½ tablespoons of salt to the other. Secure the lid, shake vigorously 60 times (Epsom salts solution) or 30 times (salt solution), and let the solutions stand. Cut circles from the cardboard to fit inside each of the saucers. Pour a thin layer of Epsom salts solution over one paper and salt solution over the other paper. Try not to pour out any of the undissolved salt. Place the saucers in a warm place and wait several days. Observe the paper daily.

Explanatory note: Long, slender, needle-shaped crystals of Epsom salts and small, white, cubic crystals of salt form on the paper.

ACTIVITY: THE CRYSTAL PAGE

Teaching note: ‘The crystal page’ on the website <http://freeweb.pdq.net/headstrong/crys.htm> details the manner in which large crystals can be produced from a variety of fairly common substances.

You will need:

- **salt**
- **Epsom salts**
- **borax**
- **alum (aluminium potassium sulfate)**
- **photographer’s hypo (sodium thiosulfate)**
- **baking soda (bicarbonate of soda)**
- **laundry soda**
- **copper sulfate.**

Acids and bases

The thought of having acid come into contact with bare skin, let alone consuming acid, sends people quivering. But this is what we do almost on a daily basis. Citrus fruits, cola and wine contain acids. Our stomachs contain hydrochloric acid, which assists in the digestion of food. All these foods have in common a sour taste (some only slightly). Bases, in contrast, have a bitter taste. Some common bases are found in soaps, toothpastes and cleaning agents.

Acids and bases can be quite dangerous to touch and can vary greatly in their strength.

Tables 6 and 7 gives the names, formulas and common uses of typical acids and bases in the community.

TABLE 6:
TYPICAL ACIDS

Acid	**Formula**	**Use or location**
Nitric	HNO_3	Used to etch copper and manufacture fertilisers
Hydrochloric	HCl	Produced in the stomach; used to clean bricks and concrete
Sulfuric	H_2SO_4	In car batteries, plastics, insecticides, detergents and pharmaceuticals
Acetic	CH_3COOH	In vinegar (used in pickling and fermentation of foods)
Citric	$C_6H_8O_7$	In lemon juice
Carbonic	H_2CO_3	In soft drinks
Lactic	$C_3H_6O_3$	Makes milk sour; builds up in muscles during strenuous exercise

TABLE 7:
TYPICAL BASES

Base	**Formula**	**Use or location**
Ammonia	NH_3	Fertilisers and detergents
Sodium hydroxide (caustic soda)	$NaOH$	Soaps and detergents
Sodium carbonate	Na_2CO_3	Manufacture of glass, washing powder and detergents
Calcium oxide (quicklime)	CaO	Bricklayers’ mortar
Lead (II) oxide	PbO	House paint (now phased out)
Calcium hydroxide (lime)	$Ca(OH)_2$	Cleaning agents
Ammonium hydroxide	NH_4OH	Milk of magnesia (for treatment of indigestion)

There are substances that can act as an acid or a base depending on the chemical environment. These are called ‘amphiprotic substances’. Table 8 provides some examples.

TABLE 8: AMPHIPROTIC SUBSTANCES

Amphiprotic substance	**Formula**
Water	H_2O
Hydrogen sulfate ion	HSO_4^-
Hydrogen carbonate ion	HCO_3^-
Hydrogen sulfide ion	HS

Properties of acids and bases

Although acids and bases can be considered different types of substances, they are intimately related. The properties of each are listed below.

Acids have the following common properties:

- They usually taste sour. This is only for weak acids. Never attempt to taste concentrated acid such as that in a car battery.
- They are corrosive. This means they break down other substances, such as metals.
- They are molecular in structure (which means their atoms are covalently bonded) and will dissolve in water to produce an electrolyte. Electrolytes are ions dissolved in water; they conduct electricity.
- They affect the colour of certain natural and synthetic dyes (called ‘indicators’). A common, natural indicator is the vegetable dye called ‘litmus’, which turns from blue to red in the presence of an acid.
- They are neutralised by bases. This means that when a base is added to an acid it loses its acidity. The addition of sufficient amounts of base can result in a neutral solution (neither an acid nor a base). The addition of a lot of base may turn an acidic solution basic.

Bases have the following common properties:

- They usually taste bitter. As with acids, never taste a strong base (e.g. oven cleaner).
- They feel slippery (they react with the natural oils in the skin and produce soap, which gives the characteristic ‘slippery feel’).
- They turn litmus from red to blue; therefore, litmus is blue in a base and red in an acid.
- They are electrolytes. That is, they are substances that conduct electricity and are ions in aqueous solutions.
- They may be corrosive. This means they break down other substances, like organic materials such as fat, hair and vegetable matter (they make good drain cleaners).
- They are generally ionic compounds.

- They are oxides (containing oxygen ions, O^{2-}) or hydroxides (containing hydroxide ions, OH^-) of metals.
- They are usually insoluble in water. A base that is soluble in water is called an ‘alkali’.

The Bronsted–Lowry theory of acids and bases

A theory about acids and bases was suggested independently by two scientists, Lowry and Bronsted, in 1923 and is still accepted as viable today. Quite simply, the theory is summarised as two statements:

- An acid is a proton donor.
- A base is a proton acceptor.

This means that, in a chemical reaction between an acid and a base, not only will bonds break, but there will also be a proton transfer from the acid to the base. The proton that the theory refers to is an H^+ ion (i.e. the nucleus of a hydrogen atom). Proton transfer does not occur in any other way. That is, it does not come from the nucleus of any other atom than hydrogen. Note that all the acids listed in Table 6 have hydrogen as a constituent of their molecules. This is a prerequisite for a substance to be an acid. However, all substances that contain hydrogen are not necessarily acids.

The following equation describes an acid–base reaction.

$$HNO_3 \text{ (l)} + H_2O \text{ (l)} \rightarrow H_3O^+ \text{ (aq)} + NO_3^- \text{ (aq)}$$

In this equation, nitric acid (HNO_3) loses a proton (H^+ ion) to water (H_2O), which acts as a base. The products of this reaction are a hydronium ion (H_3O^+) and a nitrate ion (NO_3^-).

In the next example, water acts as an acid (it is an amphiprotic substance) and ammonia (NH_3) is the base, producing ammonium ions (NH_4^+) and hydroxide ions (OH^-).

$$NH_3 \text{ (g)} + H_2O \text{ (l)} \rightarrow NH_4^+ \text{ (aq)} + OH^- \text{ (aq)}$$

Note that in these equations ‘l’ means ‘liquid’, ‘aq’ means ‘aqueous’ and ‘g’ means ‘gaseous’.

Strength of acids and bases: the pH scale

The strength of an acid is a measure of its tendency to donate a proton (H^+ ion) to a base. Alternatively, the strength of a base is a measure of its tendency to accept a proton from an acid. Because aqueous solutions (solutions containing water) of acids and bases are most commonly used, it is convenient to use an acid’s tendency to donate a proton to water, or a base’s tendency to accept a proton from water, as a measure of its strength.

In aqueous solutions, acids produce hydronium (H_3O^+) ions and bases produce hydroxide (OH^-) ions. However, as water is amphiprotic (acts as an acid and a base) it produces both H_3O^+ and OH^- ions in solution.

In pure water the concentration of hydronium (H_3O^+) ions and hydroxide (OH^-) ions remains constant (the concentration of the ions is also very small). This is because the following reaction reaches equilibrium.

$$H_2O\ (l) + H_2O\ (l) \leftrightarrow H_3O^+\ (aq) + OH^-\ (aq)$$

However, if the concentration of H_3O^+ ions increases (because the acid is in solution), the concentration of OH^- ions drops (the extra H_3O^+ ions react with the OH^- ions to produce H_2O again). Conversely, if the concentration of OH^- ions increases (because the base is in solution), the concentration of H_3O^+ ions decreases.

A scale based on the concentration of H_3O^+ ions has been devised to determine the strength of an acid or base. It is called the ‘pH scale’. The scale extends from –1 to 14. To define an acid, base or neutral solution:

- Acid solutions have a pH less than 7.
- Neutral solutions have a pH equal to 7.
- Base solutions have a pH greater than 7.

A first glance at the pH scale doesn’t make sense. You would think that, as acids in solution produce the greatest concentration of H_3O^+ ions, the pH scale for acids should have the highest values. However, because of the mathematics involved in calculating pH, it works out that the greater the concentration of H_3O^+ ions, the smaller the pH value. Therefore, the strongest acids have pH values closer to –1 and the strongest bases have pH values closer to 14.

It must be noted that some acids have a greater strength than others. This means that if a strong and a weak acid have the same concentration in a solution (same amount per volume of water), the strong acid will produce more H_3O^+ ions than the weak acid. The strong acid will produce a lower pH value than the weak acid. Conversely, a strong base will produce more OH^- ions than a weak base and so will produce a higher pH value in solution than the weak base.

The pH value for a particular acid or base depends on how much there is in solution. That is, the pH value of an acid or base depends on its concentration in solution. For example, a strong acid, such as hydrogen chloride (HCl), can be concentrated (i.e. there is a lot of HCl per volume of H_2O) in solution and have a pH value of –1 or be dilute in solution (i.e. there is little HCl per volume of H_2O) and have a pH value of 3. Therefore, a strong acid or base can be concentrated or dilute, and a weak acid or base can be concentrated or dilute.

Tables 9 and 10 give the pH values for common substances found in the home and inside you.

TABLE 9:
pH VALUES

Substance	**pH value or range**
Human gastric juices	0.9–3.0
Lemon juice	2.1
Vinegar	2.9
Orange juice	3.0
Soft drinks	3.0
Celery juice	3.8
Tomatoes	4.2
Bananas	4.6
Black coffee	5.0
Detergent	5.5
Carpet cleaner	5.9
Toothpaste	6.8
Milk	6.9
Egg white	7.6–9.5
Baking soda in water	8.4
Floor cleaner	10.0
Household ammonia	11.9
Oven cleaner	12.5

(Smith & Dwyer 1991, p. 389)

TABLE 10:
pH OF HUMAN
BODY FLUIDS

Body fluid	**Normal pH range**
Pancreatic juice	8.0–8.9
Liver bile	7.4–8.0
Blood	7.35–7.45
Fluid inside cells	7.15–7.25
Tears	7.0–7.4
Saliva	6.4–7.0
Sweat	4.5–7.5
Urine	4.5–7.5
Gastric juices	0.9–3.0

Measuring pH indicators

Some dyes (these are usually weak acids or bases) have the important characteristic that they are a different colour in an acidic solution than they are in a basic solution. Such dyes are used to indicate the presence of an acid or a base and are therefore described as ‘indicators’.

The best known natural dye that changes colour in acids and bases is called ‘litmus’, an extract from lichen. Litmus turns red in an acid and blue in a base. There are several other natural dyes that can be obtained from plants. These include extract of red cabbage, orange juice, berry juice (blackberry, blackcurrant, strawberry and raspberry), beetroot juice, black tea and black coffee. The activities listed below will allow you to experiment with such natural indicators.

Indicators can also determine the approximate pH value of an acid or base as they change colour within a small range of pH values. One such indicator, called ‘universal indicator’ (this is a mixture of several indicators) is used often in the laboratory as it changes to a range of colours depending on the pH value of the solution. For example, universal indicator is deep red in strong acid, yellow in weak acid, green in neutral solution, blue–green in weak base and blue–violet in strong base.

Neutralisation

The discussions above highlight the fact that acids, bases and neutral solutions are related to each other. A solution type is only dependent on the concentration of hydronium ions (H_3O^+) in it. For example, if you add an acid to a base this will result in more H_3O^+ ions in solution and so the pH value of the solution will drop. By adding just the right amount of acid to a base you can achieve a pH value of 7. This means the mixture is now a neutral solution. Adding more acid will result in the solution becoming acidic. Adding an acid to a base, or vice versa, is called ‘neutralisation’.

The neutralisation reaction of an acid to a base usually results in the production of an ionic compound (called a ‘salt’) and water. The water is produced by the attachment of H_3O^+ and OH^- ions, and the salt comes from other ions within the acid and base. One application of a neutralisation reaction is in the use of antacid tablets. Although the hydrochloric acid (HCl) in the gastric juices within our stomachs helps break down food, if too much acid is produced, indigestion may result. Antacid tablets contain bases, such as magnesium hydroxide ($Mg(OH)_2$), aluminium hydroxide ($Al(OH)_3$) or sodium bicarbonate ($NaHCO_3$), which neutralise the acid in the stomach.

Using an antacid to relieve an acidic problem does not apply in situations where acids or bases are spilled on the skin or are swallowed. The best remedy is to dilute the solution by adding water. Adding a base or an acid may cause more injury as it may not be the correct concentration to neutralise the acid or base, or it may not come in direct contact with the acid or base and so may cause injury on its own.

Reactions involving acids and/or bases

Acids and bases are very important to us in many ways, particularly in industry. In this topic, we will not deal with all of these processes (there are so many), but we will mention one pollution effect, called ‘acid rain’, which has resulted from industrialisation. Answers to questions about acid rain (such as ‘What is it?’, ‘How is it produced?’ and ‘Why is it dangerous to the environment?’) can be found on Environment Canada’s ‘Acid rain’ website <http://www.ns.ec.gc.ca/aeb/ssd/acid/acidfaq.html>.

The following activities relate to acids, bases and indicators found in the home; therefore, the materials are relatively safe to use and readily obtainable. As you proceed through the activities, keep in mind the processes (discussed above) that underlie each of the observations made.

ACTIVITY: CABBAGE JUICE INDICATOR

Key idea: This recipe will produce a solution that will indicate the presence of an acid, a base or a neutral solution.

You will need:

- **a strainer**
- **a jar with a lid**
- **a saucepan**
- **an uncooked purple cabbage.**

Fill the saucepan with cabbage leaves that have been torn into small pieces. Fill the saucepan three-quarters full with water. Heat the mixture to boiling; boil for approximately 1 min and then let the mixture stand until it returns to room temperature. Pour the cabbage mixture through the strainer into the jar. Store the cabbage juice in a refrigerator until needed. Test your cabbage indicator on various solutions (these are listed below). It will turn red or pink in the presence of an acid and green in the presence of a base. It will remain purple or blue in the presence of a neutral solution. If your tap water is slightly alkaline (basic) this will cause your cabbage juice to go slightly green–blue. If this occurs, boil a new batch of cabbage using bottled water.

Instead of boiling the cabbage to obtain cabbage juice indicator, you can make use of a blender and a metal strainer. Place wedges of cabbage into a blender and strain the juice through the metal strainer lined with a coffee filter or cheesecloth.

The list of household chemicals in the table below can be tested with the cabbage juice indicator. The quantities should be enough for the whole class.

TABLE: HOUSEHOLD CHEMICALS

Chemical	**Nature**	**Preparation**
Vinegar	Acid	Combine 500 mL of vinegar with 500 mL of water
Aspirin	Acid	Dissolve 40 crushed aspirin in 2 L of tap or bottled water
Lemon juice	Acid	Use full strength
Drain cleaner	Base	Add half a teaspoon of drain cleaner to 2 L of tap water (drain cleaner is extremely caustic; avoid contact with the skin)
Baking soda	Base	Dissolve 30 teaspoons of baking soda in 2 L of tap water
Antacid	Base	Dissolve 40 crushed antacid tablets in 2 L of tap water
Alcohol	Neutral	Use full-strength alcohol
Water	Neutral	Use tap water unless your local water is alkaline; if so, use bottled water
Salt water	Neutral	Dissolve several teaspoons of table salt in 2 L of tap or bottled water

Cabbage juice can be used to make a paper indicator. This may be a good option instead of having many containers of liquid indicator that can easily be contaminated.

You will need:

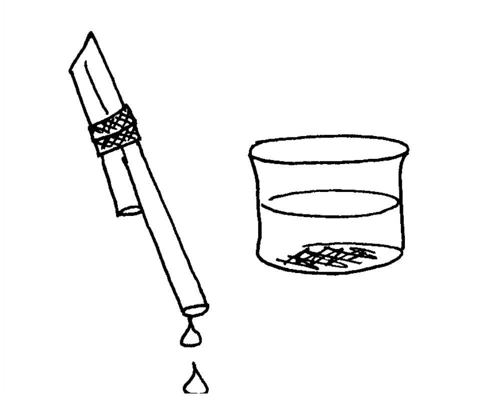
- **a clip-lock bag**
- **cabbage juice (from the recipe above)**
- **an oven tray**
- **coffee filters**
- **a bowl**
- **scissors.**

Pour 1 cup of cabbage juice into the bowl. Put a piece of filter paper in the cabbage juice and then place it on the oven tray. Continue wetting filter papers and placing them on the tray until the tray is covered with filter paper. Allow the paper to dry—this allows the water to evaporate, leaving a blue chemical on the paper that changes colour when touched to an acid (red) or a base (green). Cut the paper into strips and place the strips in the clip-lock bag for storage. These can be used to test for an acid or base. Test them out on a variety of solutions, such as ammonia, vinegar, lemon juice, orange juice, shampoos and cleaning products.

Explanatory note: In a classroom situation it is best to have the household chemicals in small containers such as baby-food jars. These will need to be suitably labelled, with each container having its own medicine dropper. The cabbage juice indicator can also be stored in a container with its own dropper.

Given that you will require a number of droppers, which may be hard to access, homemade droppers can be made out of drinking straws and tape. Cut a straw in half. Then fold over one-third of the straw and secure with some clear tape. See the figure below.

To use your drinking straw dropper, squeeze the folded part as you would squeeze the rubber part of the medicine dropper.



Teaching note: When testing the chemicals, students will require their own laboratory benches. These can be in the form of styrofoam egg cartons or, if they are not available, you could use white plastic paint trays from an art store. Students should compare a number of chemicals at once and compare the colour changes that occur.

When adding the cabbage juice indicator to the household chemicals, it is important that the dropper does not get into contact with the chemical; otherwise the cabbage juice indicator may be contaminated. Alternatively, make up indicator paper. If, after some use, your cabbage juice turns greenish or slightly pink, it is contaminated. If this happens, add a little acid to remove the green colour or a little base to remove the pink colour; alternatively, discard the juice and make a new batch.

ACTIVITY:
USING
CABBAGE
JUICE
INDICATOR

Teaching note: The following sequence of tasks is taken from Barber (1989a). Refer to the original source for details of the classroom procedure. Each task links with the next to form a lesson sequence where each new task explores deeper aspects of acids and bases.

Task 1: Cabbage juice tests

This first activity explores differences in household chemicals. There should be no mention of ‘acid’, ‘base’ or ‘neutral’ solution at this stage. Test each of the household chemicals that were listed in the previous activity. You may wish to make up a worksheet where students use coloured pencils to indicate their observations. Students should group the chemicals according to the observations from the colour changes (or lack of colour change) of the cabbage juice.

Task 2: Acid–base discovery

‘Acid’, ‘base’ and ‘neutral’ are introduced as scientific terms for the various household chemicals. Cabbage juice is identified as an acid–base indicator. Discuss the nature of each of the chemicals from the previous activity.

To undertake this activity, the students require their egg carton test solutions from the previous activity. Some will be pink, some green and some purple. To get the students to understand that acid, neutral and base solutions lie on a continuum, set them the following task. Ask them to predict what chemicals will change all of the solutions in the egg carton green. When they’ve succeeded in turning everything green, they should predict what chemicals will change all of the green solutions pink.

Which chemical was most effective at turning the liquids green? Which was the best at tuning them pink?

Did you notice any colour changes during the change from pink to green? (The solution turns purple before it turns green.)

By undertaking this activity, students can understand that the acid–base classification occurs along a continuum, rather than as separate or discrete categories. This concept can be also represented by colouring a line, going from very pink at one end to very green at the other. The colour in between is purple.

Task 3: Concentrating on amounts

This activity shows that acids and bases can vary in strength. This can be observed by the shade of colour given by the cabbage juice indicator.

You need an egg carton. In the first compartment, put five drops of vinegar (a strong acid); in the second, put four drops of vinegar and one drop of water; in the third, put three drops of vinegar and three drops of water; and so on. Test each solution with equal amounts of cabbage juice. Note the colour changes and the various shades of pink produced. Repeat the whole experiment using aspirin (a weak acid). Repeat again using drain cleaner (a strong base) and antacid (a weak base). The colour changes should indicate that some chemicals are strong acids or bases and others are weak. In addition, acids or bases can be made weaker by dilution with water.

To get an understanding: acid + base = neutral

Put five drops of cabbage juice into each of three compartments of the egg carton. Put ten drops of drain cleaner into each of these same compartments. Now count how many drops of aspirin solution it takes to neutralise (change to purple) the drain cleaner in the first compartment. Find the number of drops of lemon juice that neutralise the second compartment. Find the number of drops of vinegar that neutralise the third compartment. Discuss which acid is most or least concentrated.

Task 4: Household mysteries

From the previous activities, students should be able to determine if an unknown substance is acidic, basic or neutral. In addition, they should be able to tell if an acid or base is strong or weak. A range of household chemicals can be tested. Students may bring them from home. However, be aware that chemicals such as toilet cleaner, oven cleaner, petrol, pesticides, herbicides, drain cleaner, battery acid and ammonia are very strong acids and bases and are too dangerous for classroom use.

ACTIVITY: ACID–BASE INDICATORS

Indicators other than purple cabbage can be made from readily obtained substances. They can be made by boiling coloured flowers, cherries, beets, grapes or blueberries. The spice turmeric will turn from yellow to red when a bit of the powder is added to a base. Some may only change colour in an acid or a base but not both. Therefore, they are known as an ‘acid indicator’ or a ‘base indicator’. Once the indicators have been made, test them out on known acids and bases.

ACTIVITY: HOMEMADE FIRE-EXTINGUISHER

Can you guess what carbon dioxide gas will do to the flame from a candle?

You will need:

- **a wide-mouthed bottle or jar (or a 250 mL beaker)**
- **3–4 teaspoons of baking soda**
- **¼ cup vinegar**
- **a birthday candle**
- **a square of cardboard that will fit inside the jar**
- **matches**
- **tongs.**

Light the candle and drip some wax onto the cardboard square. Blow out the candle and position it in the hot wax until it stands alone. Place the candle on its base inside the jar. Spread the baking soda around the base of the candle. Light the candle again, using tongs to hold the match. Carefully pour the vinegar down the side of the jar. What happens?

The reaction between vinegar (weak acetic acid solution in water, $HC_2H_3O_2$) and baking soda (sodium bicarbonate, $NaHCO_3$) produces carbon dioxide gas (CO_2). The candle cannot continue to burn in a carbon dioxide atmosphere because fire needs oxygen to burn, so the flame is extinguished! This activity is a variation of the common baking soda–vinegar volcano.

Some old fire-extinguishers had a small bottle of acid inside the extinguisher but at the top. When the fire-extinguisher was turned upside down, the acid flowed into the water, which contained sodium bicarbonate. The reaction produced carbon dioxide gas, which was not flammable itself but, more importantly, greatly increased the pressure in the extinguisher. This resulted in the water spewing from the extinguisher at a great rate (even when placed vertically again).

ACTIVITY: THE ‘BAD’ TASTE OF ORANGE JUICE

Does orange juice taste awfully bitter after you brush your teeth? If so, you have a taste gene on your tongue that allows you to detect certain bitter compounds. Approximately two-thirds of the population have this gene.

When someone who has the gene brushes their teeth with toothpaste that contains sodium lauryl sulfate (SLS), they will notice this bitterness. SLS reduces the sweet taste of sucrose (sugar) and at the same time strengthens the bitterness of citric acid (responsible for the sour and bitter taste of orange juice) by approximately ten times!

If you would like to see if you inherited this gene, select a toothpaste that contains SLS in the list of ingredients. Take a sip of orange juice and note the relative strengths of the sweet, sour and bitter tastes. Rinse your mouth with water, then vigorously brush your teeth with the toothpaste. Rinse with water again, then taste the orange juice again. Are the relative intensities of the tastes very different now?

Taste begins with an ion or molecule docking in receptors on the tongue or palate. The substances that trigger sweet and bitter tastes are usually large, complex organic molecules that fit these receptors like keys in a lock. In contrast, salty and sour tastes are triggered by tiny positive ions. SLS is one of the most widely used detergent molecules. It is a large organic molecule found in toothpaste, laundry detergents and specialty detergents such as Woolite. The reason why some of you won’t notice the taste effect of SLS is because you may be insensitive to the bitter tastes of compounds called ‘phenylthiourea’ and ‘propylthiouracil’, and less

sensitive to bitter flavours such as caffeine, potassium chloride and certain preservatives. If this is the case, you have not inherited a gene from your parents that would make you sensitive to bitter tastes. Some people inherit the gene from just one parent, and they experience the bitterness effect to a lesser degree. This activity was sourced from DeCristofaro 1995, pp. 14–15.

ACTIVITY: INDIGO IMPRINTS

Key idea: In this experiment, we will make imprints of objects and then colour them 'chemically' to a beautiful blue–purple (indigo) shade.

You will need:

- **a 3 inch square piece of architect paper**
- **any solid object to imprint (a key, coins, letters cut out of paper etc.)**
- **an empty, clean peanut butter jar with its lid**
- **½ cup household ammonia**
- **several small rocks or pebbles to cover the bottom of the jar to approximately 2 inches in height**
- **a bright light (such as a desk lamp).**

Place your solid object on the yellow side of the architect paper. Let it sit under a bright light for approximately 5 min. While waiting, place the rocks in the jar and add the ammonia. Make sure the liquid is just below the surface of the rocks. Cover the jar tightly with the lid.

Remove your solid object from the light and notice the 'imprint' it leaves on the paper. Open the jar and carefully place the paper inside, taking care not to let the paper touch the ammonia. Re-cover tightly and observe the paper for approximately 5–10 min.

The ammonia fumes will turn the imprint a deep blue–purple colour, leaving the rest of the paper blank. The light chemically alters the unprotected surfaces of the paper so that it will no longer react with the ammonia, which is a chemical base. Architect paper is a very light-sensitive material, but your object has protected and preserved a small portion of it.

When you are finished, pour the ammonia down the drain and clean the jar and rocks with warm soapy water.

Take care not to inhale the ammonia fumes at any point during your experiment. This experiment was sourced from Parratore (1994, p. 76).

Oxidation and reduction

One of the largest groups of chemical reactions involves the processes of oxidation and reduction. As these processes are intimately related, the chemical reactions are called 'redox' reactions ('red', reduction; 'ox', oxidation). Redox reactions are important in our daily lives. Redox reactions occur in:

- the digestion of foods and within the cells of our bodies
- photosynthesis in plants
- decay processes, which include the formation of coal and petroleum, and fermentation processes in the production of alcohol, bread, tea, cocoa and chocolate

- the formation of air pollutants such as sulfur dioxide, nitrogen oxide and carbon monoxide
- the combustion of fuels and explosions
- batteries in clocks, radios and calculators
- metal extraction for ores
- the corrosion of metals
- the production of materials needed for making plastics, insecticides and drugs
- the treatment of pools against bacteria build up
- the bleaching and yellowing of old newspapers
- ‘setting’ and ‘perming’ hair
- alcohol breath testing by police

Redox reactions are chemical reactions and so, on the microscopic scale, chemical bonds are broken in the reactants and re-formed in the products. A special characteristic of redox reactions, in addition to bond breaking or formation, is the transfer of electrons from one of the reactants (called the ‘reductant’) to another reactant (called the ‘oxidant’).

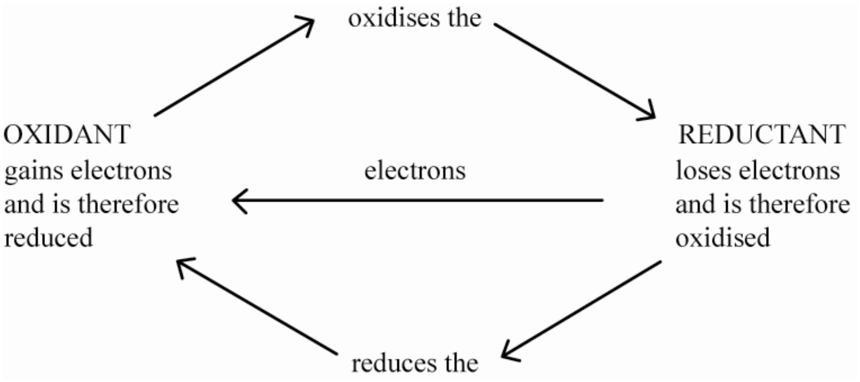
Redox reactions

Redox reactions involve the transfer of electrons from one reactant to another.

- The reactant that donates electrons is called the ‘reductant’.
- The reactant that receives electrons is called the ‘oxidant’.

In redox reactions, the two separate processes of oxidation and reduction occur simultaneously. The reactant that loses electrons (undergoes oxidation) is called the ‘reductant’ or ‘reducing agent’ because it transfers electrons to another substance and causes it to be reduced. The reactant that gains electrons (undergoes reduction) is called the ‘oxidant’ or ‘oxidising agent’ because it accepts electrons from the reductant and causes it to be oxidised. See Figure 3.

FIGURE 3: RELATIONSHIPS IN A REDOX REACTION



We will briefly look at three types of redox reactions:

- corrosion
- combustion
- electrochemical processes.

Corrosion

Corrosion is the process by which many metals react with gases in the atmosphere (sometimes with water) causing them to deteriorate. The products of corrosion have different properties from those of the metal and so their formation can be undesirable. For example, the rust formed when iron corrodes is brittle, porous, more bulky and less strong than the metal and flakes off easily. However, not all examples of corrosion are destructive. When metals such as aluminium (Al) and chromium (Cr) corrode a thin layer (called an ‘oxide’) forms over the surface stopping major corrosion. Because of this property, chromium is alloyed with iron to form stainless steel, which corrodes very little.

When metals corrode, the main corrosion product is called an ‘oxide’ of the metal. In the examples of iron, aluminium and chromium, their oxides are iron III oxide (Fe_2O_3), aluminium oxide (Al_2O_3) and chromium oxide (Cr_2O_3), respectively. Originally, the corrosion process in the formation of oxides was called ‘oxidation’ (the metal gained oxygen). However, the ‘transfer of electron’ definition has superseded this definition of oxidation.

To explain how the transfer of electrons occurs, consider the corrosion of magnesium metal. When magnesium (Mg) reacts with oxygen (O_2) in the air, magnesium oxide (MgO) is formed. The equation can be written as:

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

The transfer is not shown in the equation as the electrons that are donated by the reductant are taken up by the oxidant. In this reaction the reductant is Mg and the oxidant is O_2. To show the transfer of electrons we need to consider ‘half equations’; each shows either oxidation or reduction:

Oxidation: $Mg\ (s) \rightarrow Mg^{2+} + 2e^{-}$

Reduction: $O_2\ (g) + 4e^{-} \rightarrow 2O^{2-}\ (s)$

These half equations show that each magnesium atom loses two electrons and each oxygen atom gains two electrons (the four electrons in the equation give two electrons to each oxygen atom in the molecule). Note that in these equations ‘s’ means ‘solid’.

Not all metals react with oxygen when in direct contact with the air—for example, precious metals such as gold (Au) and silver (Ag). Iron will corrode in contact with the air but the amount of corrosion is greatly enhanced by the presence of water. In this situation the redox reaction become a little complicated.

In this reaction iron (Fe) loses electrons as part of the oxidation process. In turn, oxygen (within the water) is reduced by reacting with water to produce hydroxide ions (OH^{-}). The product of the reaction is iron II hydroxide ($Fe(OH)_2$), which is a green substance sometimes found in wet rusted pipes. Iron II hydroxide is very reactive with oxygen in the air and produces iron III hydroxide ($Fe(OH)_3$). This substance may be reinterpreted as iron III oxide with water molecules attached. As this substance dries out it loses its water molecules and goes through colour changes from black to yellow to familiar brown rust (iron III oxide, Fe_2O_3).

ACTIVITY: RUST—THE CRUST THAT MAKES STUFF BUST

Key ideas: Iron rusts easily. Steel wool has iron in it, so steel wool can rust. Besides iron, two other substances are needed to make steel wool rust. By doing this activity, you can get a good idea of what they are!

You will need:

- **a steel wool pad**
- **two clear plastic cups**
- **water**
- **vegetable oil.**

Wash all the soap out of the steel wool pad, then separate the steel wool into two equal clumps. Place each clump into the bottom of its own plastic cup.

Add water to one cup so that some of the steel wool is sticking out above the surface of the water.

Add water to the other cup so that it completely covers the steel wool. Next, carefully add vegetable oil on top of the water so that there is a layer of oil approximately 1–2 cm thick on the water's surface. Let the cups sit overnight.

Look at the cups the next day. What do you observe? In which cup did the steel wool seem the most rusted? Based on this experiment, what two substances do you think are needed to make steel wool rust?

Explanatory note: Rust is a combination of iron, water and oxygen. In the first cup, the steel wool was exposed to water and oxygen from the air. In the second cup, the steel wool was not able to come in contact with as much oxygen as it was under water and was covered with a coating of oil. Because it could not get as much oxygen from the air, there should have been less rust in the second cup. This activity was sourced from Wondernet's webpage 'Rust—the crust that makes stuff bust' <http://www.acs.org/portal/a/c/s/1/wondernetdisplay.html?DOC=wondernet\activities\metals\rust.html> (viewed 24 January 2005).

ACTIVITY: HOW MUCH AIR IS NEEDED TO RUST STEEL WOOL?

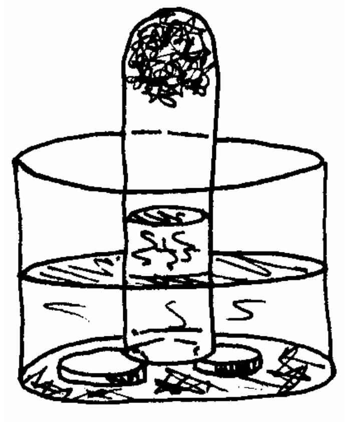
Key idea: To investigate how much air is used up in the rusting process, complete the following experiment.

You will need:

- **a large test tube or narrow glass container**
- **steel wool**
- **a large jar or beaker.**

Push the dampened steel wool into the base of the test tube so that it will remain there when the tube is inverted. Now invert the test tube into a large beaker of water. The test tube needs to be raised off the bottom of the beaker (coins resting on the bottom will work; see the figure below). Mark the level of the water. Now note the level of the water over the next 2–3 days. What do you observe? Why?

Explanatory note: After approximately 2 days, the steel wool shows signs of rust, which means it has reacted with the oxygen in the air. As the oxygen combines with the iron it is no longer in the air. As a consequence, the water rises to replace the lost oxygen. To convince yourself that something has been added to the steel wool, accurately weigh the wool before and after it rusts.



Combustion

Combustion is a redox process where a substance is chemically combined with oxygen usually with the production of heat energy. Consider the examples of the combustion of methane (CH_4, natural gas), butane (C_8H_{18}, in petrol) and candle wax ($C_{25}H_{52}$). In each of the equations below, oxygen is the reductant, receiving electrons from the other substance. The products of the reaction are carbon dioxide, water and heat energy (the reactions are therefore exothermic).

$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (g) + \text{heat energy}$$

$$2C_8H_{18}\ (l) + 25O_2\ (g) \rightarrow 16CO_2\ (g) + 18H_2O\ (g) + \text{heat energy}$$

$$C_{25}H_{52}\ (s) + 38O_2\ (g) \rightarrow 25CO_2\ (g) + 26H_2O\ (g) + \text{heat energy}$$

An important combustion process for animals is the combustion of sugars in the carbohydrates we eat. The reactions below show the combustion of an important sugar called ‘glucose’ ($C_6H_{12}O_6$):

$$C_6H_{12}O_6\ (aq) + 6O_2\ (g) \rightarrow 6H_2O\ (l) + 6CO_2\ (g) + \text{heat energy}$$

Glucose provides energy to support the internal body temperature necessary for the chemical reactions that take place within the body. This reaction also provides energy for exercise and, because of the involvement of oxygen, it explains why we breathe heavily when exercising.

ACTIVITY:
THE BURNING
SUGAR TRICK

You will need:

- **matches**
- **a piece of wire**
- **sugar cubes**
- **a candle**
- **an ashtray or fireproof mat.**

Place a sugar cube on the mat or tray and challenge another person to set it on fire using a match. They will not be able to do it.

When they give up in frustration rub some fine ash onto the cube, then set it on fire. It will burn with a blue flame. The other person may argue that it is the ash that is burning, but it is not. The ash acts as a catalyst in the combustion of sugar and is not burned itself.

Fasten a sugar cube to a wire loop, which serves as a handle. Apply ash to one corner. Hold the corner in a candle flame until it starts to burn, then remove it. The cube will drip a sticky black liquid containing carbon, and at the same time will shoot out tiny rings of smoke.

As a research exercise, find the formula for sugar and try to determine the products of the combustion reaction.

ACTIVITY: THE ART OF BLEACHING

People use liquid laundry bleach to remove unwanted colour (i.e. stains) from clothes. This bleach is a 5% solution of sodium hypochlorite in water. It also removes colour from other materials. We will use it to produce some interesting effects. Care must be taken when handling bleach.

You will need:

- **a bottle of laundry bleach**
- **coloured construction paper**
- **spoons**
- **brushes**
- **cotton swabs**
- **steel wool**
- **drinking glasses.**

The colours of construction paper quickly disappear with the application of bleach. The trick is to apply and spread bleach in a manner that will result in an artistic pattern. Pour a small amount of bleach into a glass and then experiment with different applicators, such as a spoon, brush or cotton swab. Spread the bleach around on the paper by folding or tilting the paper or by blowing through a drinking straw. A little bit of bleach goes a long way, and you‘ll be able to see the patterns almost immediately. Let your work of art dry before hanging it up for all to see.

In this case, the bleaching action can also be called an ‘oxidising’ reaction. You can prove that oxygen is made from bleach by putting two small balls of steel wool (of the same size) into two different glasses. Cover them with equal amounts of water. Add a tablespoon of vinegar to each glass, and then add a tablespoon of bleach to just one of them. After approximately 30 min the steel wool without bleach should be unchanged, but the ball with the bleach should be very rusty. Rust is iron (from the steel wool) that has combined with oxygen in the presence of water. While iron rusts easily, it happens very quickly here because the bleach is producing so much oxygen.

Safety notes: Do not leave students unattended while working with bleach. Do not let the bleach come into contact with skin or eyes; if it does, flush immediately with large amounts of water. Thoroughly clean or dispose of all materials that come in contact with bleach. This activity was sourced from Herbert 1980, p. 45.

Electrochemical processes

Electric currents involve the movement of electric charges, such as electrons and ions, within conductors, such as metals and aqueous solutions containing ions (or electrolytes). As redox reactions involve the transfer of electrons, it is possible to use redox reactions to produce electricity. To do this, one needs to physically separate the reduction and oxidation parts of the reaction. This is achieved in a device called a ‘galvanic cell’ (sometimes also called a ‘voltaic’ or ‘electrochemical’ cell).

Electrochemical cell

To understand how an electrochemical cell operates, we need to consider what happens in a redox reaction. We will take an example that was originally used to produce electricity for telegraphs and telephones. In this redox reaction, the metal zinc (Zn) is oxidised in an aqueous solution of copper II sulfate ($CuSO_4$).

When a strip of zinc is dipped into a solution of copper II sulfate it begins to dissolve. At the same time, copper metal falls out of solution and forms sediment at the bottom of the container as well as forming on the zinc metal. The zinc gets oxidised giving electrons to copper ions in the solution. On receiving the electrons, the copper ions become copper metal, which falls out of solution.

The full reaction equation and the half equations are shown below:

$Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}(aq) + Cu\ (s)$

Oxidation: $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^{-}$

Reduction: $Cu^{2+}\ (aq) + 2e^{-} \rightarrow Cu\ (s)$

The three stages of the redox reaction are shown in Figure 4. In part A, when zinc is added to the solution, a spontaneous redox reaction begins. In part B, as the zinc dissolves, copper ions are reduced to copper metal and the original blue colour of the solution begins to fade. In part C, if the zinc remains in the solution for an extended period of time, the blue colouring of the solution disappears; this means that all the copper ions have reduced to copper metal.

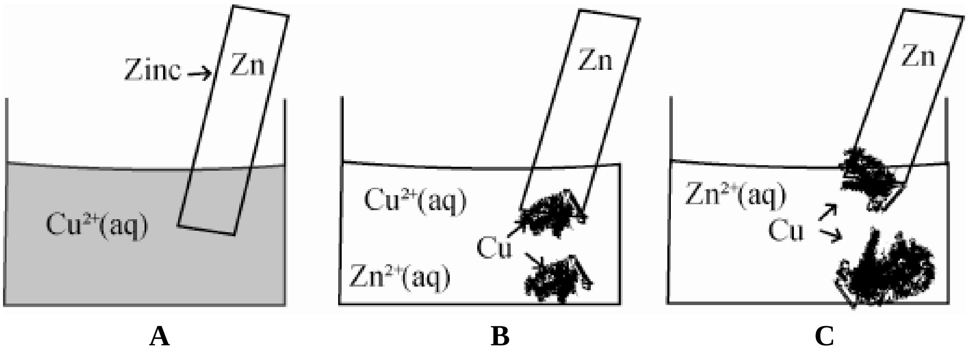
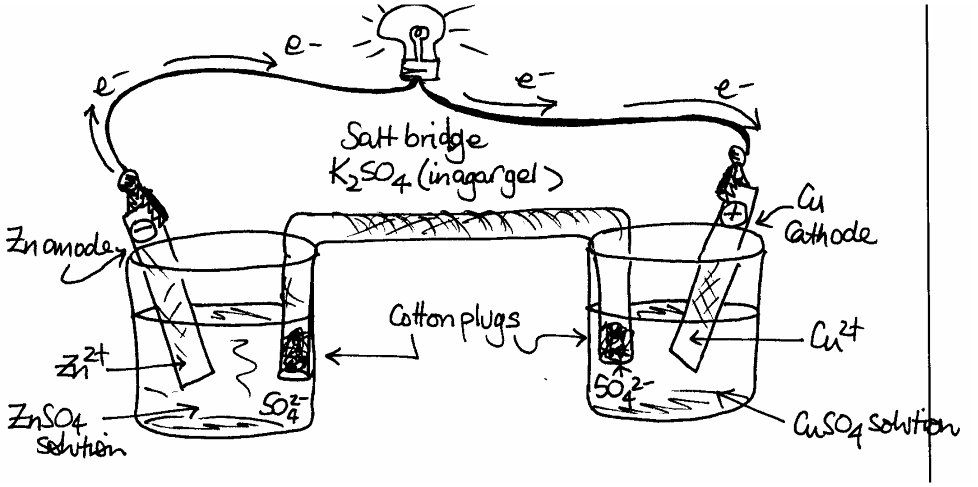


FIGURE 4: REDOX REACTION OF ZINC AND COPPER IONS IN A SOLUTION OF COPPER II SULFATE

In the example above, oxidation and reduction occur at the same place (in the one container). These processes can be separated by having two containers or half-cells. See Figure 5, which shows an electrochemical cell known as a ‘Daniell cell’.

FIGURE 5: DANIELL CELL



The Daniell cell consists of the following:

- two half-cells (containers)
- a zinc metal strip (called an ‘anode’) and a solution of zinc sulfate ($ZnSO_4$) in one half-cell
- a copper metal strip (called a ‘cathode’) and a solution of copper II sulfate ($CuSO_4$) in the other half-cell; the zinc and copper strips are also called ‘electrodes’
- a metal conductor (wire), joining the anode and the cathode, connected to a bulb, which detects the presence of an electric current
- a salt bridge, which connects the two half-cells and contains potassium sulfate (K_2SO_4).

The redox reaction is described as follows. Oxidation occurs in the cell containing the zinc. The zinc metal gives up electrons that move along the wire to the cathode and also dissolves as zinc ions into the solution of zinc sulfate. The half equation is the same as above. The only difference is that the electrons go in one direction and the zinc ions go in the other (solution).

Oxidation: $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^-$

The electrons run along the wire (creating an electric current and lighting the bulb) down through the cathode to the copper ions in the solution of copper II sulfate. Here the copper ions accept the electrons and reduce to copper metal at the copper electrode. The half equation for this reduction process is the same as above, but reduction occurs quite separately to oxidation.

Reduction: $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$

The zinc electrode is considered the anode and is negative, as this is where the electrons come from. The copper electrode is considered the cathode because it uses up electrons. The terms ‘cathode’ and ‘anode’ are used quite differently in other areas of science. For example, the cathode ray tube of a television set has a negative cathode that produces electrons that stream towards a positive anode.

The function of the salt bridge is to allow ions to move from one container to the other to neutralise any excess of charge that builds up in each container. For example, at the zinc electrode, positive charges (Zn^{2+}) fall into the solution so the salt bridge provides extra sulfate ions to electrically neutralise this excess positive charge. Similarly, in the other container, there is an excess of negative charges because the copper ions are reduced. To neutralise the solution, sulfate ions move along the salt bridge to the other container.

Batteries and cells

Cells can combine to produce greater electrical effects, such as greater voltage or current. Multiple cells that are connected are called ‘batteries’. However, common language also refers to single cells as batteries.

There a several different types of batteries each with their own oxidant and reductant. The redox reactions in many of these batteries are quite complex and will not be considered in this topic. However, the basic processes in each battery are the same. Oxidation and reduction are physically separated. The transfer of electrons occurs along a metal conductor in the form of an electric current.

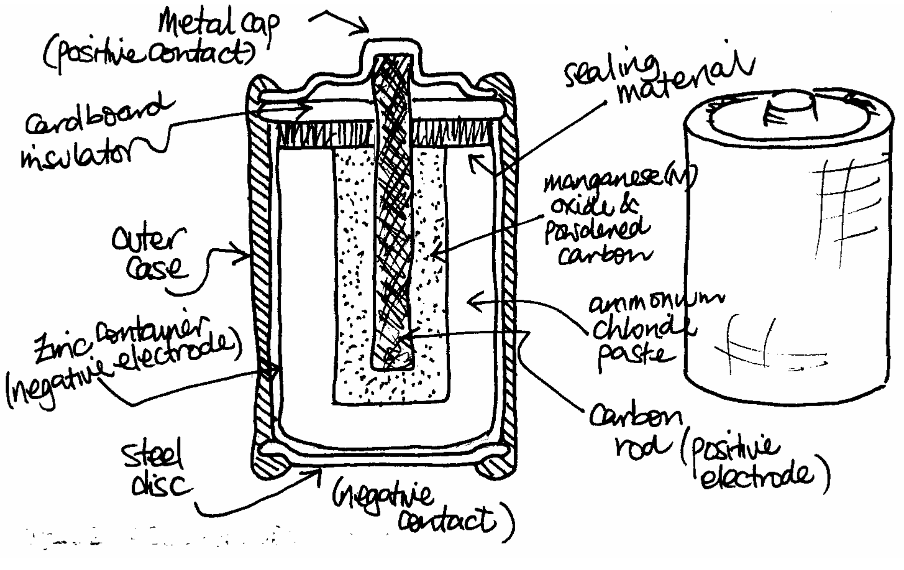
A common battery used in radios and clocks is called a ‘dry cell’. A cross-section of a dry cell is shown in Figure 6.

A dry cell consists of a zinc container filled with an electrolyte paste (which contains ions). The paste contains manganese (IV) oxide (MnO_2), zinc chloride ($ZnCl_2$), ammonium chloride (NH_4Cl) and water. A carbon rod is embedded in the paste and forms the cathode. The zinc container forms the anode.

The oxidation reaction is straightforward as the zinc metal gives off electrons and forms zinc ions that remain in the paste. The thick paste prevents the contents mixing and so a salt bridge is not required. The electrons run around a conductor and enter the battery via a carbon cathode. The electrons reduce ammonium ions in a complex reaction that also involves manganese dioxide. The redox reaction continues until the electrodes are clogged with the products of the redox reaction.

In time the zinc is eaten away. Even with a steel outer casing, the battery may then start to leak. The liquid that leaks is very corrosive to any metals and so old batteries should be removed and thrown away immediately after they have run down.

FIGURE 6: BATTERY COMPONENTS



Some of the different types of batteries in use include the following (you may wish to investigate these):

- alkaline zinc–manganese dioxide cells
- alkaline zinc–mercury cells
- lithium batteries
- lead–acid accumulators
- nickel–cadmium rechargeable cells
- fuel cells.

ACTIVITY: PUT A SHINE ON IT

This activity is taken from the 'Science is fun' website <http://scifun.chem.wisc.edu/>.

If you have any objects made from silver or plated with silver, you know that the bright, shiny surface of silver gradually darkens and becomes less shiny. This happens because silver undergoes a chemical reaction with sulfur-containing substances in the air. You can use chemistry to reverse the tarnishing reaction and make the silver shiny again.

You will need:

- **a tarnished piece of silver**
- **a pan or dish large enough to completely immerse the silver**
- **aluminium foil to cover the bottom of the pan**
- **enough water to fill the pan**
- **a vessel in which to heat the water**
- **hot pads or kitchen mitts with which to handle the heated water vessel**
- **baking soda—approximately one cup per gallon (4 L) of water.**

Line the bottom of the pan with aluminium foil. Set the silver object on top of the aluminium foil. Make sure the silver touches the aluminium.

Heat the water to boiling. Remove it from the heat and place it in a sink. To the hot water, add approximately 1 cup of baking soda for each 4 L of water. (If you need only 2 L of water, use half a cup of baking soda.) The mixture will froth a bit and may spill over; this is why you put it in the sink.

Pour the hot baking soda and water mixture into the pan, and completely cover the silver.

Almost immediately, the tarnish will begin to disappear. If the silver is only lightly tarnished, all of the tarnish will disappear within several minutes. If the silver is badly tarnished, you may need to reheat the baking soda and water mixture, and give the silver several treatments to remove all of the tarnish.

(Source: < http://scifun.chem.wisc.edu/homeexpts/homeexpts.html, viewed 10 March 2005.)

Explanatory note: When silver tarnishes, it combines with sulfur and forms silver sulfide. Silver sulfide is black. When a thin coating of silver sulfide forms on the surface of silver, it darkens the silver. The silver can be returned to its former lustre by removing the silver sulfide coating from the surface.

There are two ways to remove the coating of silver sulfide. One way is to remove the silver sulfide from the surface. The other is to reverse the chemical reaction and turn silver sulfide back into silver. In the first method, some silver is removed in the process of polishing. In the second, the silver remains in place. Polishes that contain an abrasive shine the silver by rubbing off the silver sulfide and some of the silver along with it. Another kind of tarnish remover dissolves the silver sulfide in a liquid. These polishes are used by dipping the silver into the liquid, or by rubbing the liquid on with a cloth and washing it off. These polishes also remove some of the silver.

The tarnish-removal method used in this experiment uses a chemical reaction to convert the silver sulfide back into silver. This does not remove any of the silver. Many metals, in addition to silver, form compounds with sulfur. Some of them have a greater affinity for sulfur than silver does. Aluminium is such a metal. In this experiment, the silver sulfide reacts with aluminium. In the reaction, sulfur atoms are transferred from silver to aluminium, freeing the silver metal and forming aluminium sulfide:

$$\underset{\text{silver sulfide}}{3Ag_2S} + \underset{\text{aluminium}}{2Al} \rightarrow \underset{\text{silver}}{6Ag} + \underset{\text{aluminium sulfide}}{Al_2S_3}$$

The reaction between silver sulfide and aluminium takes place when the two are in contact while they are immersed in a baking soda solution. The reaction is faster when the solution is warm. The solution carries the sulfur from the silver to the aluminium. The aluminium sulfide may adhere to the aluminium foil, or it may form tiny, pale yellow flakes in the bottom of the pan. The silver and aluminium must be in contact with each other, because a small electric current flows between them during the reaction. This type of reaction, which involves an electric current, is called an ‘electrochemical reaction’. Reactions of this type are used in batteries to produce electricity.

ACTIVITY: ALUMINIUM TRIVIA

Let’s explain why chewing aluminium foil can really hurt some people but not others. The difference is because some of us have silver fillings in our teeth. It turns out that aluminium atoms lose their electrons very easily. In the presence of our mildly acidic saliva, which acts as a catalyst, we have what amounts to a crude electric battery. Electrons flow from the aluminium to the silver amalgam filling. The electric current is felt by the nerves of our teeth and causes a downright unpleasant zing!

Aluminium foil will begin to decompose in the presence of many other acidic substances in a process called ‘oxidation’. Acids like to oxidise obliging metals. Some common acidic foods include tomato sauce, which has a pH of 3.8 (7 is neutral), or cola soda, which is even more acidic with a pH of 2.7. Tell any cooks you know never to wrap a meat loaf glazed with tomato sauce in aluminium foil for storage. After several hours the result of this contact is a greyish-black, disgusting mush of aluminium oxide.

Why doesn‘t a full aluminium cola can dissolve? The inside of the can is coated with a harmless but effective protective surface made up of long molecules called ‘polymers’ (i.e. a plastic coating). This information was sourced from Adams 1984, p. 121.

ACTIVITY:
LEMON
BATTERY

You can make electricity with a lemon. This activity was sourced from Mandell (1990, p. 58).

You will need:

- **a lemon**
- **a galvanometer**
- **two stiff copper wires**
- **a large paperclip**
- **scissors.**

If there is any insulation on the ends of the wire, the teacher should strip it off. Untwist the paperclip and twist the end of one of the wires around it. Squeeze and roll the lemon to loosen the pulp inside. Make two small cuts in the skin of the lemon 1 inch or so apart. Insert the bare wire and the paperclip through these cuts into the juicy part. The two wires should be close but not touching. Connect the free ends of the two wires to the terminals of the meter.

Now watch the meter move! What‘s happening here? Electrochemistry causes the two different metals (the copper wire and the iron paperclip) in the acid (lemon juice) to draw electrons away from one wire towards the other. They flow out from the lemon through one wire, go through the meter and then enter the lemon by the other wire.

ACTIVITY:
MAKE YOUR
OWN
GALVANOMETER

Make your own galvanometer (an instrument designed to detect electrical currents).

You will need:

- **a compass**
- **approximately 4.5 m of bell wire (from a hardware store)**
- **a small rectangular cardboard tray.**

Place the compass in the cardboard tray. Scrape off approximately 1.25 cm of insulation from each end of the bell wire. Starting approximately 15 cm from one end, wind the wire tightly around the box, circling it at least two dozen times. Leave approximately 15 cm of wire free on the other end of the tray. Rest your homemade galvanometer on the table horizontally and turn it until the compass needle is parallel to the coil of wire. Use these bell wire ends to attach to your lemon and paperclip.

ACTIVITY:
FOOD
BATTERIES

Teaching note: This activity was adapted from the MadSci Network’s webpage ‘Food batteries experiment’ <http://www.madsci.org/experiments/archive/889917606.Ch.html> (viewed 14 June 2001).

Key idea: Create a battery from common foodstuffs, sufficient to light a small light bulb or LED display.

You will need:

- **one large potato or lemon**
- **a zinc electrode—a 3.0 cm × 0.5 cm piece of zinc metal will suffice (inquire at a local hardware store)**
- **a copper electrode—a similarly sized piece of copper metal**

- **copper wire—a sufficient length of wire to create a circuit from the zinc electrode to a light bulb (or other device) and copper electrode**
- **a small light bulb—flashlight or penlight bulbs work best (experiment with other devices such as LED displays or time pieces).**

Safety: If no copper electrode is used, hydrogen gas is given off as a by-product of the reactions taking place. Be wary of performing the experiment near heat sources or an open flame. Although the voltages and amperages given off are low, care should be taken in handling the wire and other parts of the circuit.

Stick the zinc electrode all the way into the potato or lemon. Place the copper electrode on the opposite side. Connect the small light bulb to the two electrodes with copper wire. Observe what happens!

Explanatory note: A common query you may encounter is ‘Can you please tell me how to measure electricity in fruits and vegetables?’ Electricity is usually measured with a multimeter, but as there is no electricity in any fruit or vegetable (in the sense that they are like little batteries), in this case a multimeter will give no results. You can use the chemical properties of certain fruits and vegetables to generate electricity. A lemon, for example, can be made to power a small electrical device because the lemon is quite acidic (for a food). This is done by inserting a piece of zinc metal and a piece of copper metal (a zinc electrode and a copper electrode) into the lemon. You can then draw electrical power from the lemon through an external circuit and do work. (A lemon cell is approximately equivalent to a single calculator battery.)

Zinc is an active metal and will react readily with acid; acid’s active ingredient is positively charged hydrogen. So a transfer of electrons takes place between the zinc and the acid; the zinc (Zn) is oxidised to Zn^{2+} and the acid (H^+) is reduced to hydrogen gas (H_2), which you can see bubbling out around the electrodes.

Oxidation: $Zn \rightarrow Zn^{++} + 2e^-$ (zinc loses two electrons)

Reduction: $2H^+ + 2e^- \rightarrow H_2$ (hydrogen ions gain electrons)

Net reaction: $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$

Of course, this will happen whether or not you have a copper electrode present, but you need the copper electrode to draw power from the lemon cell; the copper helps channel the electrons through the external circuit. This sort of cell will work for any fruit or vegetable with some acid content; lemons are best simply because they‘re more acidic than any other food.

Further comments: Try the effect with different fruits and vegetables. How well do other citrus fruits or tomatoes work? If using potatoes, how does the size of the fruit or vegetable relate to how long the bulb stays lit? Does the pH of the vegetable relate to the amount of electricity generated? Finally, what is the maximum watt light bulb you can light from your food battery? Try using some electricity equations to calculate parameters such as resistance, voltage and current.

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