

EDUP1004 Science Foundations - Chemistry

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5 topics, in a “tutorial” style course:

1. Classification of Matter
2. Bonding
3. Properties of Materials OR How Things Break
4. Bubbles
5. Chemical Reactions

The overall aim of this course is to understand where the different properties of matter come from and how to predict them. Given that matter is made up of particles, it must be interactions between particles that are responsible for the properties of a substance. We can redefine chemistry then to be the study of the interactions between particles, and in particular, between atoms and molecules.

The Structure of each topic is effectively the same, that is,

1. An introductory paragraph
2. A longer more general description, covering the key points
3. A list of demonstrations/examples/slides and their purpose (NB the demonstrations are all simple and SAFE, I also reuse some of the demonstrations to illustrate more than one topic)
4. A list of learning objectives, that is, what I expect you to have learnt and to be able to do for each topic
5. Some sample short answer questions (I will give you answers later in the semester)
6. Some things you might like to explore with your students, including some useful websites. There are experiments/discussions you could do all the way from K to 6. The key to these experiments is to try to get your students to think about what might happen before you do the experiment.

I will also provide you, later in the semester, with a web-based Mock Exam, using the same format as your first semester exam.

All of this information, together with the notes for each topic and copies of my slides will be available on the web. In the first instance you can get in from my website: www.chem.usyd.edu.au/~mjtj but I will endeavour to have it available through Blackboard as well. The answers to the example questions and the mock exam will also be put up on the web.

I would really value your feedback and I'll put a feedback questionnaire on the website too (or you can just grab me or email me).

In particular, if you have any questions, anytime, please don't hesitate to ask!

Topic 1 Classification of Matter

Chemistry is the study of matter. Matter is anything that has mass and occupies space. We can classify matter according to its fundamental structure, what it looks like at a “microscopic” level, or its physical properties.

Structural Classification:

All matter is made up of **atoms**. An **element** is made up of a single kind of atom. Each element is given a **symbol** that stands for a single atom of that element. The atoms in an element can exist essentially independently of each other, for example in a metal, or they can be bound together, in very specific ways, into molecules.

A **molecule** is made when two or more atoms are joined together by **chemical bonds**. We can use the symbols for the atoms to describe the nature of the bonding, for example H_2 implies that two atoms of hydrogen are bound together by a chemical bond, that is, H–H, into a hydrogen molecule. The element hydrogen exists as molecules of H_2 gas. Elements can have different structural forms, for example, the element carbon can exist as diamond, graphite or as fullerenes, eg C_{60} (“buckyballs”). The different structural forms of an element are called **allotropes**.

Compounds are molecules that contain more than one type of atom **bound** together in fixed ratios. For example, the compound water consists of molecules of water H_2O , each with two atoms of hydrogen and one atom of oxygen with chemical bonds between the oxygen and the hydrogen atoms. The properties of elements and compounds depend on their structure: diamond has very different properties to graphite, water has very different properties to hydrogen or oxygen gas. A **chemical change** involves breaking a chemical bond.

Elements and/or compounds can be physically mixed together, without breaking any chemical bonds, to form **mixtures**. This is a **physical change**. Two common kinds of mixtures are **solutions** and **alloys**. Solutions are formed when something, the **solute**, dissolves in something else, the **solvent**. Water is probably the most useful solvent we have. Alloys are formed when different metals are mixed together. The properties of a mixture depend on the particular mix of components. Mixtures can be separated by using the different physical properties of their constituents.

Property Classification:

Matter can also be classified according to its physical properties. Some useful properties include melting point, boiling point, density, viscosity, conductivity to heat and electricity, strength, elasticity... Using melting and boiling points allows us to classify materials as solids, liquids or gases, these are the **phases** of the substance.

- **Solids** are not compressible, they have a fixed volume. They have a fixed shape (“edges”) and do not conform to the shape of their container.
- **Liquids** are not compressible. They have a fixed volume but no fixed shape. They “flow” to fill a container. Liquids do have an “edge”.
- **Gases** can be compressed. They have no edges, no fixed shape and no fixed volume.

Sometimes it can be difficult to tell the phase of a substance, for example gels or glasses or very dense gases.

Changing the phase of a substance, by heating or cooling, is a **physical change**.

Solids, liquids and gases can also be distinguished on a microscopic level. The particles in a solid are stationary and often very ordered. In a liquid the particles move randomly near one another and in a gas the particles move randomly throughout their container. Solids are ordered and gases are disordered. All **particles are always moving**. **Temperature** is a measure of the average speed of the particles.

Other classifications of matter are also possible. We often use the classification **metal** vs **non-metal**, or, for solids, metal, ceramic and polymer.

Demonstrations

Demonstration/Example	Purpose
Picture of gold atoms forming the word IBM, picture of iron atoms in a ring with a single electron caught inside	To illustrate that although atoms are extraordinarily small we can "see" them if we use an <i>electron microscope</i>
Examples of elements in jars, balloon filled with hydrogen	To illustrate that you can see elements and that elements can occur as solids, liquids and gases
Diamond, Lead Pencil, Soot	To illustrate the different allotropic forms of the element Carbon and to discuss their different structures
Hydrogen balloon	To illustrate that hydrogen occurs as a molecule with two hydrogen atoms joined together and to introduce the idea of a molecular formula and a structural formula
Plastic balls stuck together with sticky tape	To illustrate that chemical bonds are strong and that it takes lots of energy to pull them apart
Beaker of water	To illustrate a compound and the molecular formula and structural formula for water
Train engine and carriage	To illustrate a compound and to reinforce that it takes energy to break chemical bonds. This example also shows that bonding can be 'directional', it can occur in some directions but not others
Water and sugar	To illustrate forming a mixture
Water and sugar vs honey	To illustrate that the properties of a mixture depend on the proportions of its constituents
Balls glued together	To illustrate a solid: solids have definite shape, they are ordered and the overall layers particles are not moving (the individual particles still jiggle about)
Balls in a beaker	To illustrate a liquid: liquids have no definite shape, they flow, but they do have an edge. The particles in a liquid are moving
Bouncy ball	To illustrate a gas: the ball can bounce anywhere, gases have no shape, they can flow and they have no edge. The particles in a gas are moving quickly
Vanillin	To illustrate that gas molecules move
Beaker of water and food Colouring	To illustrate that liquid molecules move
Beakers of hot and cold water and food colouring	To illustrate that molecules move faster as temperature increases, that is, temperature is a measure of the average speed of a molecule
Flower and rubber hose and liquid nitrogen	To illustrate that we can change the properties of a substance by cooling it down

Learning Objectives

At the end of this topic you should know:

- the meaning of, and distinction between *atoms, elements, molecules, compounds and mixtures*
- the meaning of the term *allotrope*
- the meaning of the terms *chemical change* and *physical change* and the distinction between them
- that chemical changes involve large amounts of energy and are often irreversible
- that physical changes involve small amounts of energy and are relatively easy to reverse
- that atoms can be represented by symbols
- how to use the periodic table to identify the names and symbols of atoms

- that compounds are represented as molecular or structural formulas using the symbols of the constituent atoms
- what a solution is and the meaning of the terms *solute* and *solvent*
- some of the physical properties used to characterise substances
- how to distinguish the different *phases* of a substance: *solid*, *liquid* and *gas* on the basis of their physical properties
- how to distinguish the different *phases* of a substance: *solid*, *liquid* and *gas* on the basis of their constituent particles
- the difference between a *crystal* and an *amorphous* solid
- what a *meniscus* is and how it forms
- that phase changes are physical changes
- that all particles are always moving
- that temperature is a measure of the average speed of particles
- that as temperature increases a substance becomes more random
- how to identify a *metal*, *non-metal*, *ceramic* and *polymer* by their properties

Questions

1. What elements have the symbols Fe, At, Li
2. What is the chemical symbol for the elements Arsenic, Helium, Phosphorus
3. Give examples of allotropes of two elements
4. What does the molecular formula $C_6H_{12}O_6$ mean
5. If molecule contains one atoms of carbon and two atoms of oxygen what is its molecular formula
6. List some solutions that we see in everyday life, identify the solute and solvent for each solution
7. List three methods that can be used to separate mixtures
8. State whether the following are physical or chemical changes. Justify your answer.
 - extracting gold dust from silt in a river
 - boiling an egg
 - boiling water
 - making toast
 - making gin and tonic
9. How do the dissolved and molten states of a substance like salt (sodium chloride) differ?
10. Name the three common states of matter. What are their characteristics?
11. Why liquids are almost incompressible but gases can be compressed easily?
12. Why do solids maintain a fixed shape?
13. What is temperature?
14. What are the physical characteristics of a metal?

Things you might like to explore with your students:

The Motion of Particles: Set up 3 beakers (any see-through container) with the same volume of ice water (have the ice just melted), room temperature water and warm/hot water. Drop a few drops of food colouring into each one and watch the colour spread throughout the beaker. In each case the colour will diffuse in the water until the water is a uniform colour. You could time this with a stopwatch and ask the students to tell you when the water is uniformly coloured. This diffusion is called “Brownian Motion” and Albert Einstein was the first person to accurately describe it mathematically. The warm/hot water will be uniformly coloured first, the room temperature water second and the ice water last. Try to get your students to think about what they might see before they do the experiment and then to think about explanations for what they do see. Ask the students to describe (or draw) how the colour moves through the container. The two key ideas are (i) that the water molecules are moving and this movement spreads the colour around (indeed movement is in all directions) and (ii) that hotter things move faster so the food colouring is dissolved faster, ie temperature is a measure of the speed of the particles. You can follow this up by asking your students if they think it’s easier

to dissolve sugar in warm water or cold water (eg in the process of making toffee), an old jar of crystallised honey works too, you can redissolve the sugar by heating the honey up. You could also set this up as a race – one team of students has warm/hot water and the other cold water – and see how long it takes them to realise that it's not a fair competition.

A very cool variation of this experiment which could be used as a teacher demonstration is the “Hot and Cold Water” experiment: <http://www.abc.net.au/science/surfingscientist/pdf/teachdemo10.pdf> – indeed there are lots more experiments on this site (and the abc one: <http://www.abc.net.au/science/>)

Phase Changes: Pour equal amounts of water into 2 clear containers, maybe Friday morning. Mark the level of the water with marker pen. Cover one of the containers with cling wrap and leave the other one open (put both in a sunny or warm spot in the classroom). On Monday (or longer if it's been cold and wet!) examine the water levels in the containers. Ask your students to explain what has happened to the missing water (it evaporates into the air as water vapour). The cling wrap might even have droplets of water on it for them to see, ask them how the droplets got there. You could extend this a little by getting them to measure the water levels every day and by asking them what factors they think will influence how fast the water evaporates. You could then encourage them to test their ideas by design new experiments.

Phase Changes, the Motion Particles: Perfume/Vanilla/Perfumed Oil: set a little dish with some liquid perfume/vanilla/oil in it in one corner of the classroom. Time how long it takes for the student sitting furthest away to smell it. Ask the students to explain how the perfume smell got to the other side of the room. Repeat this on another occasion (or with a different smell) when the perfume/vanilla/oil is *heated* (over some warm water or over a candle), again time how long it takes for the person sitting furthest away to smell it. Try to get your students to think about what might happen before they do the experiment. Here the liquid or solid is evaporating/subliming to make gas particles which will then diffuse through the air (ie it is not the liquid that is diffusing). Again, if you heat the particles they move faster, so temperature is a measure of the speed of the particles.

Making Solutions: Salt or Sugar – Which Dissolves Fastest in Water? You could set this up as an experiment in class. Take 3 clean, dry clear plastic cups. Using a permanent marker label one “water”, the second “sugar” and the third “salt”. Put half a teaspoon of ordinary granulated sugar into the cup labelled sugar and half a teaspoon of table salt into the cup labelled salt. Using a measuring cup, put one cup (250mL) of water into the “water” cup. This cup serves as a “control”, that is, there is nothing in it so it is a comparison to judge when the sugar and the salt have dissolved. Start a stopwatch as one cup of room temperature water is poured into the “sugar” cup (you could also use a clock with a second hand) and either start a second stopwatch when a cup of water is poured into the “salt” cup, or pour in the water when, say, the stopwatch reaches 1 minute. Time how long it takes for the sugar and salt to completely dissolve. **Do not stir the cups.** If you have a number of students doing this, collect all the times and work out an average time for the sugar and the salt to dissolve (again you could set this up as a race). Get your students to suggest explanations for what they observe. The salt should dissolve faster because salt grains are normally much smaller than sugar grains. Because they are smaller, the salt grains will “fit” in the spaces between the water molecules faster than the sugar will (there's a bit more to this though...). You could lead the class discussion by getting them to visually compare the sugar and the salt grains. Once your students have suggested an explanation you could test it. For example compare icing sugar, castor sugar and granulated sugar (or grind the granulated sugar into smaller particles in a mortar and pestle). You can also get them to think about what would happen with cold or warm water, what would happen with stirring and what might happen with different liquids eg vinegar, lemonade, windex... (and then do the experiments!) or maybe ask them what would happen if you used a sugar cube...

Separating the Components in a Mixture: The ink in a black water soluble texta (and even a permanent marker) is a mixture of a number of different dyes (different colours). You'll need a black texta and some filter paper (eg coffee filter paper) or chromatography paper (from a science supplier) or blotting paper (art shop) or even paper towel, in fact paper towel works fine. Cut the paper into a long strip maybe 3cm by 12 cm and draw a line with black texta 3 or 4cm above the bottom of the strip. Half fill a glass with water. Place the paper strip in a glass of water so that only the end is in the water and the texta line is **above** the water level. Secure the paper above the glass (eg stickytape it to a pencil over the top of the glass). For the thicker papers you can bend the paper over the rim of the glass and secure it with a rubber band (so the paper doesn't fall in the water and ideally doesn't touch the side of the glass). The water will be drawn up slowly into the paper by capillary action (see topic 4) and, as it moves through the paper, it carries different sized dye particles different distances (ie the big heavy ones wont move as far). After about 24 hours you'll see bands of colour on the paper indicating the different pigments making “black”. You could repeat this with different colour makers, different types of marker and different solvents, you might even discover a solvent that will remove permanent marker!! Separating things like this is called “chromatography”.

Topic 2 Bonding

Why are atoms different? Why do elements have different properties? To answer these questions we need to understand more about the structure of atoms.

The Structure of Atoms

Atoms are made up of 3 basic particles: **protons**, **neutrons**, and **electrons**.

Particle	Symbol	Charge	Relative Mass
Proton	p+	+1	≈ 1
Neutron	n ⁰	0	≈ 1
Electron	e-	-1	≈ 1/1837

Most of an atom is empty space. Most of the mass of the atom is concentrated in the **nucleus** of the atom. The nucleus contains protons and neutrons and is held together by the strong nuclear force. Electrons are held in the atom by electrostatic interactions with the nucleus **UNLIKE CHARGES ATTRACT** (*in fact almost all of chemistry can be explained by the idea that unlike charges attract*). These forces are very large → electrons move rapidly through the available atomic volume, sometimes at speeds close to the speed of light.

Atomic Number, Isotopes

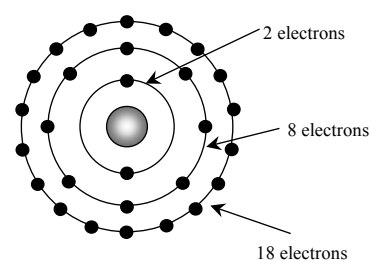
Atoms are often designated using the notation ${}^A_Z X$ where **A** = the mass number = # p + # n and **Z** = the atomic number = # p. The number of neutrons in the nucleus of a given atom is thus (A-Z). The properties of an atom are determined by its atomic number, Z.

Isotopes have the same atomic number but have different numbers of neutrons and therefore, different mass numbers. They have effectively identical chemical properties because Z is the same. For example, hydrogen has 3 isotopes:

name	symbol	composition of nucleus	natural abundance
hydrogen	${}^1_1\text{H}$	1 proton, 0 neutrons	99.985%
deuterium	${}^2_1\text{H}$	1 proton, 1 neutron	0.015%
tritium	${}^3_1\text{H}$	1 proton, 2 neutrons	0%

Electrons

A neutral atom has the same number of protons and electrons, eg a C atom has atomic number Z=6 so it has 6 protons and a neutral C atom must have 6 electrons. Positively charged atoms, **cations**, have more protons than electrons, a Na⁺ cation has 11 protons and 10 electrons. Negatively charged atoms, **anions**, have more electrons than protons, an O²⁻ anion has 8 protons and 10 electrons. Electrons do not collapse into the nucleus because they are spinning around it in special orbits (just like the Earth doesn't collapse into the sun and a bicycle doesn't fall over when you're riding it). These orbits, or "shells" get bigger, and can hold more electrons, the further out from the nucleus you go. The first shell can only hold 2 electrons and the second can hold a maximum of 8 electrons. The third shell also holds 8 electrons, but can hold up to 18.



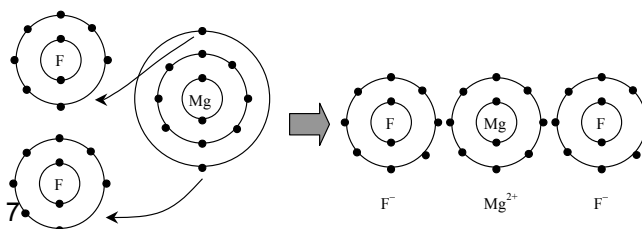
The Formation of Ions

Only the noble gases helium, neon, argon etc (group 18 on the far right-hand edge of the periodic table) are normally found in nature as atoms, all other elements are normally found as networks of atoms, either joined to themselves (eg O₂, Au metal) or other elements (eg CO₂, stainless steel) **WHY?** Phenomenologically (ie as a phenomenon, without explanation), electrons like to form closed shells with particular numbers. These "magic numbers" are 2, 8, 8, 18... corresponding to "full" outer shells of electron.

Elements that do not have the right number of electrons to make a filled shell like to "**share**" electrons with each other so that they **do** have a filled shell. Eg:

F = Ne - 1e- wants to 'borrow' 1e- and form an F⁻ negatively charged ion, or **anion**

Mg = Ne + 2e- wants to 'lend' 2e- and form an Mg²⁺ positively charged ion, or **cation**



If Mg and F get together, the “happiest” (most stable) conformation is for Mg to lend an electron to 2 F atoms, ie, MgF_2 is the most stable network between F and Mg atoms.

Elements on the left hand side of the periodic table need to lose electrons to obtain a full outer shell, they form cations easily and are called **metals** (the “lost” electron conducts heat and electricity). Elements on the right hand side of the periodic table need to gain electrons to form a full outer shell, they form anions and are called **non-metals** (they don’t conduct).

If the electrons are “transferred” from the one atom (eg Mg) to another (eg F’s) then we say that the bond between Mg and F is an **ionic bond**. The bond forms because ions of opposite charge attract each other: **UNLIKE CHARGES ATTRACT** (*in fact almost all of chemistry can be explained by the idea that unlike charges attract*)

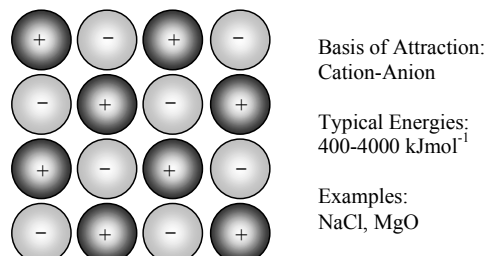
Ionic Solids

Ionic bonds tend to be *extremely* strong. The strength of the interaction depends on the charge on the ions and how large they are. The greater the charge, or the smaller the internuclear distance, the stronger the ionic bond.

The characteristics of ionic solids are that they

- (i) are hard, rigid, crystalline, and have high melting points (sometimes thousands of degrees!); *to break or melt the solid you need to break the extremely strong ionic bonds*, the ions are held very tightly together in a crystal lattice.
- (ii) are brittle; *ionic solids crack rather than bend when struck with enough force*; as the crystal is distorted the ions no longer line up and repulsions between like charges crack the solid apart
- (iii) do not conduct electricity as solids; *electrons are held too tightly by the ions* and
- (iv) do conduct electricity as molten liquids, or in solution; *the ions themselves are released to carry the current*.

The crystal structure of an ionic compound is a three dimensional lattice through space (e.g. cubic, tetrahedral, octahedral). The actual structure depends on how big the ions are and the relative numbers of positive and negative ions.

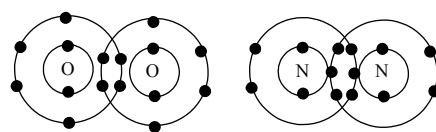
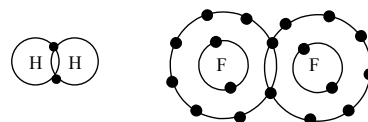


Covalent Bonds

In a **covalent bond** an electron from *each* atom is shared, *i.e.* covalent bonds involve electron *pairs*. In this way, *both* atoms think they have a filled shell, eg 2 H atoms can share their electrons so that each atom has 2 electrons or 2 F atoms can share two electrons so that the outer shell of each atom has 8 electrons (ie we attain the ‘magic’ numbers). The H–H or F–F bonds are called **covalent bonds**. Bonds form because atoms are more stable (have lower energy) in molecules than they do by themselves (with the exception of the noble gases: He, Ne, Ar, Kr, Xe and Rn which already have full outer electron shells)

Definition: A covalent bond is a pair of shared electrons.

Some atoms can share more than one pair of electrons at a time. Such sharing results in a **double bond** (for 4 electrons, or two *pairs* of electrons), eg O_2 or a **triple bond** (for sharing 6 electrons, or three *pairs* of electrons), eg N_2 . Double bonds are stronger (more stable) than single bonds. Triple bonds are stronger than double bonds.

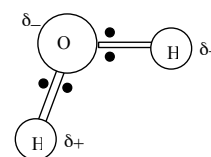


The atoms must be very close together for electrons to be effectively shared, and typically, only the smallest atoms (notably first row, but also S, and very rarely Si) can form double or triple bonds.

The electrons that are involved in bonding must be in the outermost shell, the **valence shell** (the inner electrons are held too tightly by the positively charged nucleus to be able to participate in chemistry, ie to be shared with or transferred to other atoms).

Polarity

Electrons in a covalent bond between two identical atoms are equally shared. However, the electrons in covalent bonds between two *different* elements are not shared equally, the bonding electrons spend more time on one of the atoms than the other. For example, the O atom in water is better at attracting electrons than the H atoms, electrons spend more time on the oxygen than the hydrogen giving the O atom a



slight negative charge and the H atom a slight positive charge. The O–H bond in water has slight ionic character and is called a **polar** bond. The δ^+ and δ^- symbols are used to indicate there is only a small fraction of charge on each atom as distinct from the full + or – charge found in ions.

Electronegativity (EN) is the ability of an atom in a molecule to attract a bonding pair of electrons towards itself. Electronegativity is highest in the top right hand corner of the periodic table and lowest in the bottom left corner. If a bond is formed between two different atoms, the electron pair of the bond will be attracted towards the atom with the higher EN. This results in the atom with the higher EN having a partial negative charge and the bond is termed **polar** (eg the O-H bond).

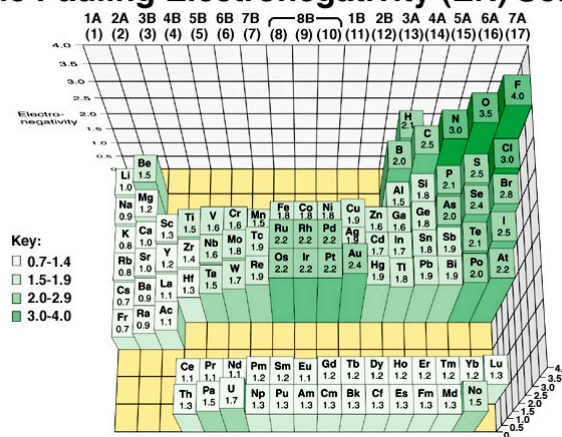
If the electronegativity difference between the atoms in a bond is *small* then the electrons in the bond are *effectively* shared by both atoms and we consider it a covalent bond (eg the C-H bond).

If the electronegativity difference between the atoms in a bond is *medium* then the more electronegative atom preferentially attracts the electrons in the bond and we consider the bond polar (eg the O-H bond).

If the electronegativity difference between the atoms in a bond is *large* then the electron is effectively transferred completely to the more electronegative atom and we consider the bond ionic (eg MgF_2).

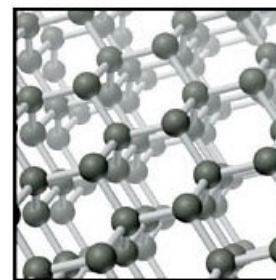
The Pauling Electronegativity (EN) Scale

Martin S. Silberberg, Chemistry: The Molecular Nature of Matter and Change, 2nd Edition. Copyright © The McGraw-Hill Companies, Inc. All rights reserved.



Properties of Covalent Compounds

Typically covalent bonds are *very* strong. The strength of a covalent bond depends on the kind of atoms involved in the bond and the type of bond formed. There is a broad inverse correlation between the strength of a bond and its length. Short bonds are stronger than long bonds. Covalent bonds are also **directional**, they occur only in certain directions, giving a molecule shape. Because of this directionality and the filled electron shells, covalently bound **network materials**, where the covalent bonding extends throughout the whole material, are typically *hard and brittle*. For example, in diamond each carbon is covalently bonded to four other carbon atoms throughout the entire crystal and in quartz (sand) each silicon atom is bonded covalently to four oxygen atoms and each oxygen atom bonds to two silicon atoms in a pattern that extends throughout the sample. In a network solid bonds need to break for the material to be deformed. As one bond breaks the entire structure is compromised so deformation results in cracks.



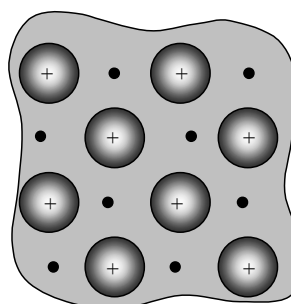
Diamond ● Carbon

Most covalent substances consist of separate molecules with weak **intermolecular** forces holding them together. In these **molecular solids**, each molecule is free to act more or less independently. These observations are borne out by:

- molecular compounds tend to have low melting and low boiling points (eg dry ice, solid CO_2)
- molecular solids are soft because the molecules can slide past each other with small stress applications, eg polymers consist of covalently bound atoms in a chain. The covalent bonding within the chains is relatively strong, whereas the forces holding the chains together are relatively weak. As a consequence, many polymers can be deformed easily since the chains can slide over each other. And
- the molecules remain intact in liquids and gases

Metallic Bonds

The outermost electrons in most **metals** are only loosely bound to their nuclei because of their relative remoteness from the positively charged nuclear cores. These electrons combine to form a “sea” of negative charge binding the positively charged nuclei together and the *metallic* electrons are shared amongst *all* the atoms in the sample. Metallic bonds are non-directional, and, as a consequence, the bonds do not break when the metal is deformed. This is one of the reasons for the high ductility of metals. Metals are also good conductors of electricity and heat because the electrons are free to move through the sample. Metals such as Cu, Al, Ag



Basis of Attraction:
Cations-delocalised electrons

Typical Energies:
 $75\text{-}1000 \text{ kJmol}^{-1}$

Examples:
Ag, Au, Al

and Au have metallic bonds. Metals tend to form crystals as the atoms pack together as closely as possible. They are therefore also dense materials.

Demonstrations

Demonstration/Example	Purpose
Periodic Table of the Elements	To show some of the information contained in the periodic table and how it can be used to predict properties of atoms
Movie	To illustrate how salt dissolves in water into its component ions, and how dissolved salt can conduct electricity – whereas a salt crystal can't
Animation	To illustrate shared electrons in a covalent bond
Toy train	To illustrate attraction between opposite 'charges'
Plastic Molecule	To illustrate that covalent bonds have a fixed direction in space
Movie	To demonstrate how electronegativity difference leads to the different kinds of bonds
Vanillin	To show that when the vanillin crystals melt and vapourise the vanillin molecule retains its integrity
Models and samples and slides	To illustrate different crystals
Copper wire and ping pong balls glued together in layers	To illustrate how malleable a metal is

Learning Objectives

At the end of this topic you should have gained the following knowledge:

- know the names and characteristics of the sub-atomic particles: *electrons*, *protons* and *neutrons*
- know the general structural features of atoms
- understand the terms *atomic number* and *isotope*
- understand that electrons form *shells* around the nucleus
- understand that atoms gain electrons to form anions or lose electrons to form cations so that their outermost electron shell is full
- understand how *ionic* bonds form as a result of electron transfer from one atom to another
- know the relationship between the strength of an ionic bond and the charges and sizes of the ions
- know why ionic compounds are brittle, have high melting points and conduct electricity when molten or dissolved
- be able to rank the melting points of simple ionic solids
- understand that a full outer electron shell can also be attained by sharing electrons in a *covalent* bond and that a covalent bond is a pair of shared electrons
- understand that double bonds are two pairs of shared electrons and triple bonds are three pairs of shared electrons
- understand that sharing may not be "equal" leading to the concept of polarity in covalent bonds which can be quantified by the *electronegativity* difference between the atoms on either end of the bond
- know that atoms in the top right hand corner of the periodic table are the most electronegative and atoms in the bottom left corner are least electronegative
- know what network and molecular solids are and what their respective properties are
- know what a metallic bond is and what the general properties of metals are

Questions

1. Where is the mass of the atom concentrated?
2. What fundamental property of the atom determines which element it will be?
3. Give the number of protons, neutrons and electrons in an atom of Fluorine.
4. Where are the electrons in the atom located?
5. Define the terms *isotope*, *atomic number*, *cation* and *anion*
6. How many electrons would be in an O^{2-} ion?
7. "Sodium chloride (salt) contains ionic bonds." Explain this statement.
8. Why doesn't solid sodium chloride conduct electricity whereas molten or dissolved sodium chloride does?
9. Why are ionic crystals so stable?
10. Which will have the higher melting point NaCl or MgO?
11. "Water molecules contain covalent bond between the H atoms and the O atom." What does this mean?
12. Why do atoms join together to form bonds?
13. In general, which is the strongest: a single bond, a double bond or a triple bond?
14. Consider the molecules H_2 , HCl and HF. Which molecule would have the most polar bond? Which molecule would have the least polar bond?
15. In the HF molecule, which atom has a slight positive charge and which a slight negative charge?
16. Why does diamond have such a high melting point (3550 °C)?
17. When liquid water (H_2O) boils or evaporates, does the gas consist of H_2O molecules, molecule fragments or separate H and O atoms? Explain.
18. List four characteristics of a solid metal.
19. When you touch a piece of metal and a piece of wood both at room temperature, which feels colder? Why?
20. Magnesium metal is easily deformed by an applied force, whereas magnesium fluoride is shattered. Why do these two solids behave so differently?

Things you might like to explore with your students:

Characteristics of a Metal: Give the students something metallic (eg a stainless steel mixing spoon) and something wooden (eg a wooden spoon). Ask them to tell you which one feels hot and which one feels cold. The metal spoon will feel cold because it conducts heat away from their hands whereas the wooden spoon feels warm because it does not conduct the heat away (it's an insulator). Metals not only conduct heat but electricity too, and wood is a "heat" (thermal) insulator and an electrical insulator. Use this as a beginning to get your students to think about other properties of metals and non-metals (eg "bendability", density – does it float or does it sink...). You could also get them to compare the properties of cling film and aluminium foil.

Growing Crystals: Crystals are best grown from a supersaturated solution, that is, one holding more than it wants to of the solute to be crystallised. To make a supersaturated solution, dissolve as much sugar, salt etc as you can in warm water and then tip the liquid (not the undissolved material, you could use a coffee filter to filter out any undissolved material) into another container then wait for the solution to cool down. As it cools, crystals will start to form. They then grow over a period of hours to weeks. Crystallisation also works well if you provide a nucleating site for the crystal to form on, for example you can use a "seed" crystal – formed by pouring a small amount of your saturated solution onto a plate, letting the drop evaporate, and scraping up the crystals formed on the bottom to use as seeds – put these seeds into the supersaturated solution to grow particularly large crystals. In a smooth container (like a beaker) you can use a piece of cotton, a knotted piece of fishing line, a metal wire, a pipecleaner (this makes great Christmas decorations if you bend it into a star shape first or make a snowflake shape)... to act as a nucleating source. Watch and wait! Different substances make different coloured and different shaped crystals. You can speed this process up (but you do get smaller crystals) by helping the evaporation process along eg by putting a teaspoon of salt water solution onto a glass dish (big surface area to evaporate) or into an afoil pie dish. You can put some black paper in first and then pour over a bit more solution, this looks great when the water has evaporated (it can be hard to see small white crystals on a white saucer). If you put salt in one solution and sugar in another, as the water evaporates you'll see different crystals and get random crystal patterns. You could get your students to draw the patterns. If the crystals are very small you might need to use a magnifying glass to tell the difference between the shapes of the salt and sugar crystals. Sugar and salt are easily obtainable, other

easy to get things that make beautiful crystals are Borax (in the cleaning section of the supermarket) and Alum (sometimes in the spices section although you might need to get it from your local high school or you can buy it in a crystal growing kit eg from Questacon) Baking Powder (sodium bicarbonate, in the baking section). To make a sucrose crystal like the one in the lectures, you need to first make some relatively large seed crystals. Then pick the largest, best shaped of these and tie some fishing line around it. Tie the other end of the fishing line to a pencil, for instance, and hang the seed crystal into a supersaturated solution of sugar so that it is covered by the liquid but does not touch the bottom or the sides of the container. Cover the container with a coffee filter paper or a paper towel – you don't want things dropping into it. It may, however, take weeks to get a really big crystal and you may need to replenish the supersaturated solution in the process.

Good crystal growing websites are: <http://www.waynesthisandthat.com/crystals.htm> and www.seawhy.com/xl.html

The RACI (The Royal Australian Chemical Institute) sponsors a crystal growing competition for schools although I could only find the 2006 details when I looked: <http://www.raci.org.au/national/chemistry/2006%20RACI%20Crystal%20Growing%20Competition.pdf> .

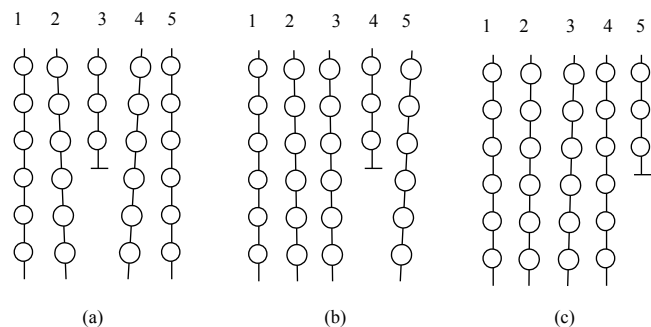
Topic 3 Properties of Materials OR How Things Break

We can understand the properties of solids in terms of their underlying structure. As liquids are cooled they ultimately form one or more stable structures, often exhibiting some kind of periodic order. The most general feature of these low temperature structures is a high density of either particle packing, for example, metals and many organic solids, or the formation of interparticle bonds, such as in network covalent solids and ceramics. It is the high density of interactions that results in typically small compressibilities and high strengths. The density of interactions is also responsible for the resistance to deformability - the essential characteristic of a rigid solid.

The simplest (and roughest) measure of the bond strength in a material is its melting point. The strength of a material will scale roughly with the melting temperature. The strength, however, will typically fall far (up to 100,000 times) below that predicted on the basis of the bond energies in a perfect crystal. Material fracture and hence material strength is actually dominated by the presence of **defects**.

[ASIDE] Many materials have different crystal structures at different temperatures. Materials that undergo a transition from one crystal structure to another are **polymorphic**, or, for an element, **allotropic**. For example CO₂ has 5 structural forms and tin has two.

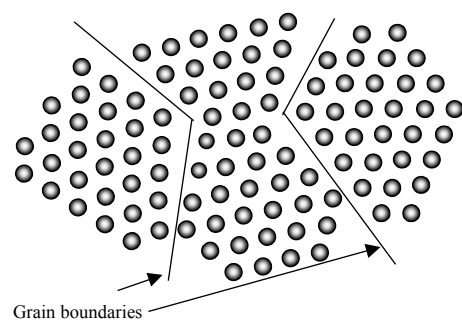
Materials deform by a process called **slip**. Within each crystallite the atoms will **dislocate** along certain preferred directions (depending on the crystal structure). The easiest dislocation to understand is the **edge dislocation**. As stress is applied to a material with an edge dislocation, originally in plane 3 (Fig (a)), a bond in plane 4 breaks and a new bond forms to the original dislocation (b). The effect of this is to move the dislocation one plane over to the right. Repeated shifts of this kind lead to incremental movements of the dislocation and eventually the formation of a step in the exterior surface (c). The material is now deformed.



In metals, the atoms are relatively close together and relatively easy to move, so slip is easy. **Metals deform easily**. In ionic solids, the positive and negative ions are often different sizes and the crystal cannot **pack** closely together. As a result dislocation movements in ionic crystals are difficult. Further, ionic and covalent interactions are large and it is difficult to break these bonds. The combination of these factors means that **ionic solids and ceramics tend to be hard and brittle**.

In practice, a given material is generally not a single crystal but rather **grains** of individual single crystallites, which range in size from nano- to millimetres. These materials are called **polycrystals**. The grains are separated by **grain boundaries**, which are less densely packed than the crystallites. Slip tends to be frustrated by grain boundaries. Polycrystalline materials tend to be much stronger (harder to break) than crystalline materials (an example is 'Corelle' dinnerware).

The ability to break something depends on defects in the material. Defects are more likely in larger samples; *it is often easier to break a large sample than a very small one!*



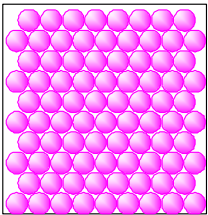
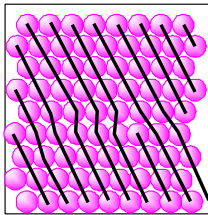
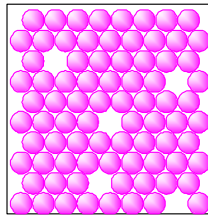
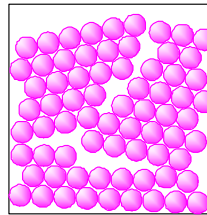
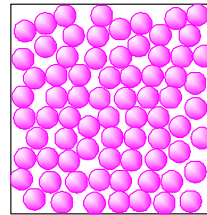
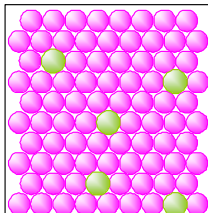
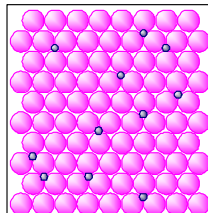
Glasses

Glasses are defined to be non-crystalline (**amorphous** – that is, without form) materials. They can be classified as solids or liquids and are often just called “glasses”. Examples of glasses include naturally occurring obsidian, and silicate glasses (windows). Glasses are formed by the rapid cooling (**quenching**) of a liquid to prevent crystallisation. Eventually glasses will transform to crystalline structures, however, this may take thousands of years. Glasses are generally perceived to be very rigid materials. Silica glass, however, behaves at room temperature as an elastic solid and can be bent as long as its breaking strength is not exceeded (for example optical fibres - a tennis ball 'bouncing' off a window). Silica glass also possesses a very large tensile strength (about five times that of steel), although tensile strength is reduced by surface defects (which are very common, especially in *cheap* glass like window panes!). For example, glass may be **cut** by scratching it with a diamond tip. Once the glass has been scored, when it is stressed, the rigidity of the glass concentrates the entire tensile force on the few interatomic bonds at the base of the crack. This causes a propagation of the crack into the glass and, since glass does not possess grain

boundaries, there is nothing to impede the propagation of the crack and the glass fractures. Glass can be “toughened” by tempering – forming a skin on the outer surface of the glass that is under compression, or by laminating glass and plastic together (to make car windscreens and bullet proof glass).

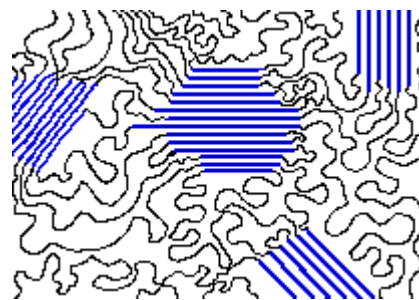
- A number of additives can be added to molten silica (molten sand) to change the properties of the glass:
- Na_2O (soda, ~15%) lowers the viscosity and melting point (from 1723° to 850°)
- CaO (lime, ~10%) makes glass insoluble in water
- Al_2O_3 (~1%) increases durability
- Both Na_2O and CaO also alter the network structure making the mixture more fluid and more likely to form a glass. “Soda-lime” glass is normal window glass and makes up about 90% of glass products.
- MgO slows down the rate at which the glass crystallizes (a glass-forming agent).
- B_2O_3 (instead of Na_2O) produces a borosilicate glass that expands less on heating (pyrex).
- PbO produces lead glasses that are ideally suited for high-quality optical glass (“crystal” glasses).

Types of Solid:

		Defects		
				
Ideal or “perfect” crystal No defects It takes more energy to deform a crystal (you have to break all the bonds in a row) and crystals are often brittle	Slip Dislocation An extra row of particles has been inserted Can also get screw dislocations	Vacancies Particles are missing from their locations in an ideal crystal	Polycrystalline Material Material consists of grains of individual crystals, separated by grain boundaries Make by annealing or “work hardening”	Glass “Amorphous” (without form) material with short range order but no long range order
		These mean the material is often easy to deform but they impede the progress of a dislocation often making the material harder to break		
“Impurities”				
				
Substitutional Solid Particles have been inserted into the crystal lattice replacing the original particles Eg alloy materials, can get solid solutions in an alloy		Interstitial Solid Particles have been inserted into the spaces in the crystal lattice Eg steel has interstitial carbon atoms in an iron crystal		
Impurities can pin a dislocation, preventing it from moving through the solid				

Polymers

Polymers can be naturally occurring, eg starches and proteins or man-made, eg plastics, with repeating units in their structure (the –“mer”). They are characterized as very long chain molecules with repeating units along the polymer backbone. These units could include chain branches or cross-links between polymer chains. Because the chains can often slide over each other, polymers are often very flexible. Areas in a polymer where the chains are closely packed in a regular way are said to be **crystalline**. Areas that are disordered are described as **amorphous**. The percentage of crystallinity in a polymer is very important in determining its properties. The more crystalline the polymer, the stronger and less flexible it becomes.



The physical properties of a polymer, such as its strength and flexibility and its degree of crystallinity depend on:

1. **Chain length** - in general, the longer the chains the stronger the polymer, both because of the intermolecular forces that sum over the entire length of the polymer and the likelihood that the polymer chains will be entangled. Polymers don't have a single chain length but a distribution of chain lengths, ie not all particles in a polymer are the same!
2. **Side groups** - polar side groups give stronger attraction between polymer chains, making the polymer stronger. They also allow interactions with polar solvents (like water) making the polymer soluble in water.
3. **Branching** - straight, unbranched chains can pack together more closely than highly branched chains, giving polymers that are more crystalline and therefore stronger. Branched chains can entangle, making the polymer more flexible.
4. **Cross-linking** - if different polymer chains are linked together extensively by covalent bonds, the polymer is harder and more difficult to melt.

For example, long, unbranched polymer chains can pack together into crystalline domains that alternate with amorphous regions (like the figure on the previous page). This gives a relatively strong, stiff material with some flexibility. Smaller branched polymers can't easily adopt crystalline structures and tend to be weaker, less dense and more easily deformed. In general properties like ductility, tensile strength and hardness rise with increasing polymer chain length, eventually leveling off.

The properties of a polymer can also be varied by co-polymerization, effectively joining different polymers together. Polymers exist in condensed phases (solids, liquids/gels or glasses rather than gases) because of their giant size, it takes more energy to break the interactions between polymer chains than it does to breakdown the chain itself.

Tin

Normal or white β tin is a metal tetragonal crystal structure density of 7.31g/cm^3 . Below about $13.2\text{ }^\circ\text{C}$, grey α tin, however, is the most stable structure. α -tin is a semiconductor with a diamond structure and a significantly lower density than β -tin, 5.77g/cm^3 . The 27% expansion of tin from white to grey causes most tin objects to crumble – tin “pest” or tin “plague” (in the middle ages it was thought it was a disease process).

- The tin buttons on Napoleon's soldiers' coats disintegrated while on their retreat from Moscow.
- Since tin pest looks like the tin has become diseased, many in the middle-ages attributed it to Satan as many tin organ pipes in churches fell victim.
- In 1910 Scott was beaten to the South Pole by Amundsen. Scott's expedition returned on foot towards caches of food and kerosene they had left on the way in but there was no kerosene, the cans, soldered with tin, were empty. The members of the expedition died only 11 miles from a massive depot of supplies

Demonstrations

Demonstration/Example	Purpose
Models and samples	To illustrate different crystals
Corn Cob	To illustrate a dislocation
Galvanized iron	To illustrate a polycrystalline material with crystal grains large enough to see
Copper Wire/Paper clip	To illustrate the properties of a metal and to demonstrate "strain hardening" – you can't get the kink out of the paper clip and eventually it will break at that location
Ripping Paper	To demonstrate that it is often easier to break something in one direction rather than another
Breaking objects	To demonstrate that it is more difficult to break a small object
Mirror, Glass Fibres	To illustrate that glass is elastic
Glass	To show how a small defect can lead to glass breaking
Broken Glass	To show the broken edge of the glass is very smooth but curvy, rather than straight (the glass broke along a defect but the break is curved because there are no crystal planes)
Different kinds of plastic	To illustrate some of the properties of polymer materials
Rubber Band	To illustrate ductility and elasticity and to show that if it is stretched too far it loses elasticity and will eventually break
Balloon and skewer	To show how polymer chains can rearrange if they have enough time to do so
Silly Putty and Tomato Sauce	To show the unusual properties of long chain polymers, that is, how their material properties can vary with the nature of the forces applied to them
Tin	To talk about different crystal structures and speculate on Napoleon's defeat outside Moscow

Learning Objectives

At the end of this topic you should have gained the following knowledge:

- understand how materials fracture and why it is easier to break a substance than would be predicted from bond strengths alone
- know why polycrystalline materials are stronger (require more energy to break) than single crystals
- know why small objects are often harder to break than large ones
- understand what a glass is and why silicate glass is easy to break
- understand the different characteristics of polymers and how they may affect the properties of a polymer material

Questions

1. What factor governs how easily a material will break?
2. Why are polycrystalline materials stronger than single crystals?
3. Why do large objects break more easily than small objects?
4. Of the first 5 listed in the "Types of Solid" section, which would require the most energy to deform?
5. Why is steel stronger than iron?
6. What are glasses?
7. How are glasses formed?
8. Why can you see through a glass window?
9. What characteristics of a polymer determine its properties?

10. Why was Napoleon defeated outside Moscow?

Things you might like to explore with your students:

Awesome Balloon Experiment: Blow up a balloon and tie it off. You might like to let out a bit of air first so that it is not overinflated. Spread cooking oil over a bamboo skewer. **Slowly** twist the skewer through the very top of the balloon (where the balloon is thickest) using gentle twisting motion. Keep pushing and twisting until the skewer emerges near the balloon tie. Pull and twist the skewer out slowly through the tied end. As the skewer goes completely through the top hole the balloon will slowly deflate. After you've pulled the skewer out jab the balloon with it and it will pop. This demonstration indicates that the polymer chains that make up the balloon can rearrange themselves about the skewer, sealing the hole, as long as you move the skewer slowly. If you move it too fast (like when you pop the balloon at the end) there isn't enough time to reseal the hole.

How Different Materials react to Different Forces: materials can react differently to pulling, stretching, bending, squashing... Give your students a range of objects eg a balloon, clay, straw, pipecleaner, pencil, plastic bag, lunch box... and get them to predict what will happen if they squash, pull, stretch, bend the object and then get them to try each (get them to draw pictures).

Silly Putty and Tomato Sauce: long chain polymers can undergo shear thickening or shear thinning, ie as you shake them they can get thicker (more viscous) or thinner (less viscous). To make silly putty: make a 1:1 mixture (by volume, eg 1 cup of each) of cornflour and water by slowly adding the water to the flour (you can colour it with food colouring if you like). Once you have a sloppy paste pour some into someone's hand to show that it flows easily. However, if you squeeze a small amount of the paste between your fingers it solidifies into a hard blob. The blob only survives as long as it's being squeezed (compressed). You see the same effect if you pour some onto your hands and clap them together. No (or very little) splash occurs because the liquid solidifies between your hands but then becomes a liquid again as the pressure is removed. With the tomato sauce the reverse happens. When you tip the bottle upside down, no liquid flows out, however when you shake the bottle liquid flows easily. In "simple" liquids, viscosity decreases as temperature increases (eg candle wax). Cornflour is made of a long chain polymer – starch. It is an excellent example of a "shear-thickening" material, it becomes almost solid under pressure. The tomato sauce, however, gets runnier the more you shake it, it is a "shear-thinning" liquid. After you stop shaking it thickens to its original viscosity.

Plasticine: Roll some plasticine into a ball. Ask your students what they think will happen to its shape if you dropped it onto a hard floor. Drop it onto the floor from shoulder height and get them to draw before and after pictures. Now ask them what they think will happen if the plasticine hit the floor with a greater force. This time throw the plasticine down at the floor and get them to look at it to see if it caused a greater change in shape.

Polymers: get your students to bring in (or find) a range of polymer materials (plastics, foams...). Get them to describe the different properties of the polymers to see that you can make something with pretty well any property you like. Get them to put together a list of things that are made from polymers (ie plastics, foams...) You might like to follow this up with getting them to think about how plastics are recycled. A good example is how polar fleece jumpers are made from recycled PET plastic bottles (this makes quite a good assignment, or project). Alternately you could discuss natural vs man-made materials.

TOPIC 4 Bubbles

The Unusual Properties of Water

Water molecules are strongly **polar** and liquid water contains networks of **hydrogen bonds** that give water its unique characteristics. For example, water expands on freezing to allow room for the geometry of the hydrogen bonds, its boiling point is higher than you would expect based on its size and it exhibits unusually high surface tension.

[ASIDE: It is often stated that life depends on these anomalous properties of water.]

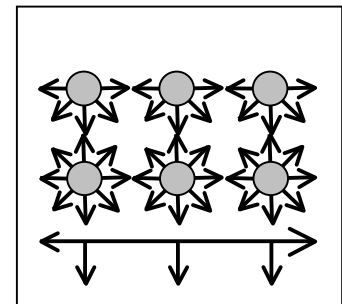
The large heat capacity and thermal conductivity helps organisms (which are mainly composed of water) to regulate their temperature. Water is also difficult to evaporate and this helps prevent dehydration and allows for “evaporative cooling” (sweating). Because water is polar it is an excellent solvent for dissolving polar and ionic compounds (salts). It can form hydrogen bonds with biological molecules like proteins that determine the 3-dimensional protein structure and the way the protein functions. Water can also ionise to form H^+ and OH^- ions, leading to **acid** and **base** chemistry.

Because water expands when it forms ice, rivers and lakes freeze from the top down so fish etc don't freeze. Furthermore the ice insulates the water below so it doesn't get too cold and the convection currents in water carry oxygen down into deep water.

A great website describing the properties of water is <http://www.lsbu.ac.uk/water/>.

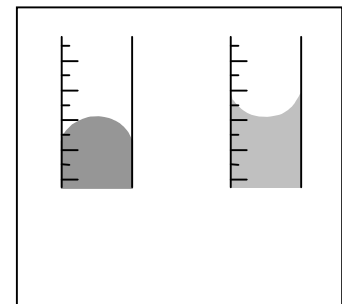
Surface Tension

All molecules attract. In particular a molecule of liquid water attracts all the surrounding water molecules and is also attracted by them. *Inside* liquid water all of these attractive forces balance out (ie the liquid is stable as a liquid). At the surface, however, water molecules are attracted by the molecules below and to the side of them but there are only air molecules above. The air is a gas and the gas molecules are on average much further away from a surface water molecule than the other water molecules in a liquid (molecules in a liquid are closer together than those in a gas). As a result the liquid water molecules below the surface yield a force pulling the surface water molecules down *into* the liquid. There is also a sideways force as water molecules on the surface attract each other. This makes the surface act like *glad wrap* (it's called a **meniscus**) as it wraps up of the liquid water. The surface force is called **surface tension**. You can see surface tension when a wet insect tries to climb out of the water: the surface tension pulls it in. Other insects, like water striders, exploit surface tension to stride or skate across the surface without sinking.



The force between water molecules is a **cohesive** force but there is also an **adhesive** force describing the adhesion of the water molecules to any other surfaces they are in contact with. When the forces of adhesion are greater than cohesion the water will **wet** the surface. If the cohesive forces are stronger the water will **refuse** the surface and form droplets. For example water wets clean glass but not dirty glass.

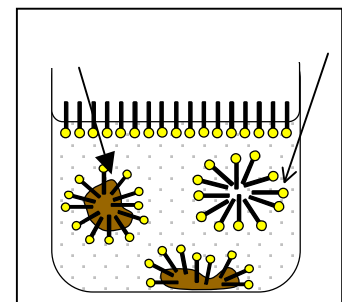
The adhesive force between water and glass is stronger than the cohesive force between water molecules. This causes water to rise up the walls of any glass container it is in. The water adheres to the glass and the water/glass surface is a concave. Mercury, however, does not wet glass (cohesion is stronger than adhesion) and forms a convex surface. In narrow glass tubes (capillaries) the adhesive forces pull water up the tube until the adhesive force is balanced by gravity. This is called **capillarity** and pulling the water up the tube is called capillary action – it's how trees obtain water.



Detergent Molecules/Surfactants

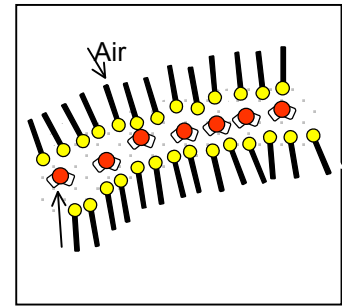
Detergent and soap molecules have two different parts: a polar or ionic part called the head group, and a long hydrocarbon tail. The head group is **hydrophilic**, ie it is attracted to water and the tail is **hydrophobic**, ie it is repelled by water (it likes fat or greasy substances).

When you put a detergent into water, the hydrophobic tails change the structure of the water and disrupt hydrogen bonding between the water molecules. As a result the hydrophobic tails are expelled to the surface of the water so the detergent molecules oriented with their hydrophilic head groups in the water and their hydrophobic tail groups away from it, pointing into the air. This is why molecules like detergents are called **surfactants**, or **surface active agents**. Replacing water molecules on the surface with detergent molecules lowers the surface tension of



the water. Indeed the ability of detergents to remove dirt is due partially to this decrease in water's surface tension: the lower surface tension makes it easier for the water to **wet** (or penetrate) the fabric.

Within liquid water, detergent molecules arrange themselves into little globules, called **micelles**, or layers, called **membranes**. In a micelle or membrane all the tails are pointing together towards the inside and all the head groups are pointing outwards towards the water. When detergents meet dirt (or oil or grease...) they surround the dirt with their tails pointing towards the dirt and their heads pointing toward the water. Agitation (eg the washing machine) encourages the head groups to carry the dirt up to the surface of the water (scum on the surface) where it can be rinsed away. Soap bubbles are formed by soapy membranes. The membranes around cells are formed by phospholipids – biological soap molecules.



Bubbles

Bubbles consist of air surrounded by films of soapy water. The detergent molecules line the interface between the air and the water. Their head groups are dissolved in the water and their tails point towards the air. The detergent molecules serve three purposes in the creation of bubbles. (a) They lower the surface tension of the water. For example, when water sprays from a tap small bubbles form, but they burst almost immediately because the high surface tension of the water pulls the water molecules in the bubble back into the main body of the water. The bubble wall gets too thin to remain intact and it quickly bursts. In soapy water the surface tension is much lower (about 1/3) so the molecules of the bubble are less stressed and the bubble can survive longer. (b) The detergent molecules are very elastic and allow the bubbles to deform without breaking. The forces between detergent molecules are much weaker than those between water molecules and this makes the surface more elastic and more able to deform. (c) The detergent molecules also slow the evaporation of the water film and so extend the life of the bubbles.

Over time the water in the soap film will migrate, under gravity, to the bottom of the film or bubble and the film at the top will become thinner and eventually burst. The bubble's life can be extended by adding substances to the water to make it thicker, or more viscous. These additives are things like glycerin and can include sugar, honey and gelatin (glycerin can be obtained from a pharmacy and is used commercially to keep products such as make-up and fruit moist). Bubbles will also pop if they touch the ground, clothing or a dry finger because the film will **wet** the surface.

Bubble Recipes: Tap water is often 'hard', containing mineral salts that can react with the soap or detergent to either remove it from your solution or change its properties. 'Soft' tap water is OK but bubbles work best with distilled water (you can buy this either at the supermarket or a hardware store or make it yourself). There are a lot of commercially available surfactants: different brands of dish washing liquids or powders, shampoos, liquid soaps, washing detergents... Liquids work best because they dissolve in water more easily.

Two common recipes are:

- Recipe A: 1 to 3 parts liquid dishwashing detergent + 6 parts water
- Recipe B: 1 to 3 parts liquid dishwashing detergent + 6 parts water + 1 to 4 parts glycerin (Too much glycerin makes the bubble mixture too heavy and prevents the formation of bubbles. You can get glycerin from a pharmacist or a cake decorating store).

You will have best results if you let the soapy solution rest for a couple of days, but if you are impatient, you can use it immediately. A cold solution makes longer lasting bubbles. To minimize evaporation make bubbles in shady areas when the air is as still as possible. They will also last longer if the air is very humid, for example after a storm. Stay in open areas so the bubbles won't run into dry objects and keep your bubble tools really wet with bubble solution to make the thickest films possible. There are lots more bubble recipes at <http://www.bubbles.org/html/solutions/formulae.htm> and various other bubble sites.

Why are bubbles round? Bubbles are "minimal surface" structures, that is, they hold the air inside them with the smallest possible surface area. The geometric form with the least surface area for any given volume is a sphere. However, when a normally round bubble is surrounded by other bubbles, it can be made to take on a seemingly odd shape (eg a cube), although, if the surrounding bubbles are popped the central bubble reverts to a sphere.

Why are bubbles coloured? White light is separated into colours as it reflects from two surfaces of a thin film (eg a bubble or an oil slick). Each colour reflects from both the outer and inner surfaces of the bubble film. Because the bubble has some thickness, this light is slightly out-of-phase and the two reflecting rays can interfere. If the interference is constructive the colour becomes more intense, if it is destructive the colour effectively cancels out and we can't see it. As the bubble gets thinner (eg under gravity) the two

reflected rays of light coincide and cancel and the bubble loses its colour. Bubbles are a cheap and very easy way to demonstrate that light is made up of colours and that colours can interfere.

Really good websites with information suitable for kids:

<http://www.exploratorium.edu/ronh/bubbles/>

<http://www.bubbles.org>

Demonstrations

Demonstration/Example	Purpose
Water bending with toward a charged rod	To illustrate that water is a polar molecule
Floating Needle or paper clip on water (An easy way to do this is to first put some tissue paper on the surface and then put the needle on top.)	With dry fingers and a dry needle, a needle will float on the surface of a glass of water. The water <u>wets</u> the tissue paper and it sinks but surface tension holds the needle or paper clip on the surface even though they are denser than water.
Floating needle + detergent (or Capillary Tube + detergent)	Adding even small amounts of detergent means that water can't rise up the capillary tube as far and the needle sinks. The detergent lowers the surface tension of the water.
Beaker, sticky tape and balls	To illustrate a <i>meniscus</i> in terms of particles sticking to each other or to the sides of their container
Capillary Tube	The adhesive forces between water and glass are stronger than the cohesive forces between water molecules and pull water up the capillary tube until the adhesive force is balanced by the weight of the water (narrower tubes can pull water higher).
Water on a leaf and a paper towel	To demonstrate "wettability". The water beads on the leaf but "wets" the paper towel
Unfolding Flower	As water is drawn into the paper by capillary action it straightens out the paper fibres and unfolds the petals
Bubble and Balloon	To demonstrate that the minimal surface for a given volume is a sphere
Adding Bubbles and Square Bubble	Bubbles stick together and the resulting minimal surfaces can be very complicated.
Coloured Bubbles	Light is made up of colours. The colours demonstrate interference patterns and you can see turbulence (in the flow of the film) as shifting patterns of colour.

Learning Objectives

At the end of this topic you should have the following knowledge:

- Know that water is a very unusual substance
- Understand what surface tension is
- Understand what cohesive and adhesive forces are
- Understand how a concave or convex meniscus forms
- Understand capillary action – why water is drawn up into narrow glass tubes and what governs how high the water can rise
- Understand the key parts of a soap/detergent molecule (a surfactant) and how they interact with water, air and grease or dirt
- Understand how the presence of detergent can change the surface tension of water

- Understand the structure of a thin soapy film
- Know how to make bubbles and why the 3 bubble ingredients are important
- Understand why bubbles are round
- Understand the factors that can make a bubble pop and how to make a bubble last as long as possible
- Understand in principle why a bubble film is coloured and that it can act like a mirror.

Questions

1. Why does mercury in a glass thermometer have a convex meniscus but water in a capillary tube a concave meniscus?
2. Will water rise higher in a thinner or thicker glass tube?
3. Why does the addition of a detergent reduce the surface tension of water?
4. How does a detergent work to clean dishes or clothes?
5. Why do soaps work better in 'soft' water than in 'hard' water?
6. What are the roles of the detergent (surfactant) in making bubbles?
7. Why do you sometimes add glycerin to a bubble recipe?
8. Do bubbles last longer on a sunny or a rainy day?
9. Why are bubbles spherical?
10. Why do bubbles pop when you catch them in your hand but remain 'living' if they land on your soapy finger?

Things you might like to explore with your students:

Making the best bubbles. Obtain a range of dishwashing detergents, shampoo, bath soap, etc. With water, make a solution in the ratio of 1 to 10 for each surfactant. In a place without wind, blow a bubble of about 7 cm in diameter using a straw and measure how long it lives before it bursts. Repeat this about 5 times for each detergent to obtain average values. The best detergent is the one that produces bubbles that last the longest. You can also ask students to design bubble recipes using various amounts of glycerin.

What tools form the best bubbles? The basic blowing tool is a straw. The basic blowing method is to stick the straw into the bubble solution, and place a finger over the top of the straw to trap some solution inside. Let most of the solution drain out, and gently blow out the rest. Other possible bubble generating tools that can be dragged through the bubble solution include plastic pipes, slotted spoons, frames from straws and string..., Have your students determine which tool produces the largest bubble.

How do bubbles mass together? Tell your students to blow one bubble on a wet table. This bubble will be the nucleus for bubble construction. Get them to add bubbles to this central nucleus and determine the maximum number of bubbles that will share sides with the central nucleus and that will adhere to each other.

Exploring Optics. As well as showing how white light is made up of colours and that light can constructively or destructively interfere, you can use this as an introduction to what colour is and how we see colours. Bubbles can also be used to study concave and convex reflections. The bubble's outer convex surface acts like a convex mirror, producing upright images. The inside back concave face of the bubble produces either upright or upside down images depending on how far the object is away from the bubble, just like a concave mirror.

Exploring charges. Bubbles possess static electrical properties. Stroking a Lucite rod with silk or an ebony rod with fur produces a static charge. Bring the rod near a bubble sitting on a wet surface. The bubble will respond by moving towards the charged rod. With some practice, you can make the bubble glide across the table toward the charged rod. Just take care to keep the rod dry.

How do you cross a membrane without bursting it? If you touch a film with a dry finger it will burst. If you will wet your finger with the soapy solution, the film will not burst and you will be able to penetrate it.

Should you rinse dishes after washing them? If you leave dishes in a drainer to dry they will be covered by a thin film of detergent when the water evaporates. Detergent manufacturers would assure you that this is harmless. Washing this film off by rinsing the dishes may not be very efficient: the detergent may like to adhere to the dishes more than to water. Vinegar in the water, however, will remove the detergent from the dishes and will also aid in removing stains and grease (the acid can react with the grease).

Why do bubbles in a bubble bath disappear when the soap falls in the water? The bubbles in bubble bath form because the surfactants in the bubble bath lower the surface tension of the water and make the surface more elastic. Different surfactants, like the bubble bath and the soap, give different elasticities. In the

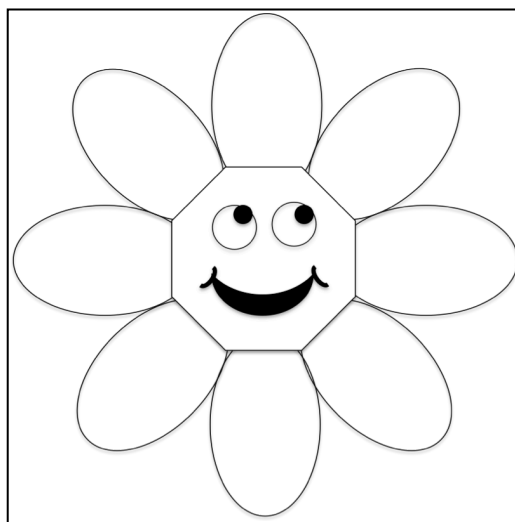
bath the poorer surfactant (the soap) is preferentially adsorbed to the surface of the water and displaces the better bubble former (the bubble bath) and the bubbles break. In this circumstance the soap is a “defoaming” agent.

Surface Tension of Water – How many drops of water fit on the surface of a five cent coin? This experiment is from http://www.abc.net.au/science/surfingscientist/pdf/lesson_plan07.pdf and is aimed at upper primary school children. Ask the class to predict how many drops of water might fit on a five cent coin. Then give them a coin, some water and an eye dropper. Count how drops you can add making sure that the drop on the end of the eye dropper **does not touch** the water on the five cent coin. Ask the class if they were surprised at the answer and then ask them to think about things that might affect the results (and even to check them). Another extension would be to ask what would happen if you added a drop of detergent.

Sandcastles: You can build sandcastles out of wet sand – water forms bridges between the sand particles (capillary action) and hold the sand together. You can't make sandcastles out of dry sand! What happens to a sandcastle as the sand dries out? How does the weather (eg the humidity/water content of the air) affect how long the sandcastle survives?

Unfolding Flower: Cut out a flower from the template. You can get your students to colour it in. Fold the flower so the petals cover up the smiley face. Float the folded flower on the water and watch. Ask your students what has happened. For example you can get them to see if the paper is wet (does the dry flower unfold by itself).

Magic Milk: put some milk in a saucer and put, say, 3 drops of different coloured food colouring in a triangle on the milk. Now put a few drops of detergent into the centre of the saucer. The milk will swirl around as the detergent assists colloidal particles of milk fat to form. As the milk swirls, the colours swirl and start to mix. You can get students to draw the results. You can ask them what they think is happening. Does the same thing happen in skim milk or low fat milk or cream? What if you dropped in some vinegar instead of detergent (see next section for what happens when you mix milk and vinegar). What happens with no detergent? (This is effectively Brownian motion, like in topic 1).

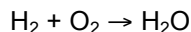


Topic 5 Chemical Reactions

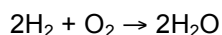
Chemical bonds are broken and made in chemical reactions. There are a number of key aspects to a chemical reaction: the overall reaction must conserve matter and energy, reactions go faster as you heat them up, if you make a change to one of the conditions then the reaction will go in such a way as to reverse that change.

Chemical Equations

We can use atomic symbols and molecular formulas as a shorthand way of writing chemical reactions, for example, hydrogen and oxygen can react to form water:



But there are more oxygen atoms on the right hand side of the reaction than on the left hand side. Because matter is conserved no atoms can be created or destroyed in a chemical reaction, so we need to **balance** the number of atoms on each side of the arrow:



that is, two molecules of hydrogen react with one molecule of oxygen to form 2 molecules of water

Energy

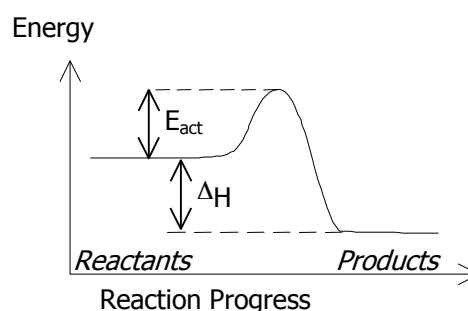
In a chemical reaction bonds between atoms in the original molecules (reactants) break and new bonds are formed with other atoms to produce different molecules (products). Reactions can occur only if the reactants collide with sufficient **energy** to allow bonds to break: *all* chemical reactions require energy to proceed because at least some of the original bonds must be broken before energy can be released through the formation of new bonds (for example, we need a match to ignite the gas in a gas stove...) We can increase the average amount of energy per molecule by heating the reactants up: reactions go faster at higher temperatures because more collisions have sufficient energy to break bonds.

*Breaking bonds requires energy, it is an **endothermic** process.*

*Making bonds releases energy, it is an **exothermic** process.*

If the new bonds made release more energy than it took to break the original bonds, there will be an overall release of energy and the reaction is exothermic. If the new bonds made release less energy than it took to break the original bonds, there will be a net requirement for energy to be supplied and the reaction is endothermic. We can measure both the energy required to initiate a reaction, its **activation energy**, E_{act} , and the overall gain or loss of energy, the **reaction enthalpy** ΔH , and plot these diagrammatically in a reaction energy diagram, as energy against some measure of reaction progress. If the energy of the products is below that of the reactants, the reaction is **exothermic**, and ΔH is negative (ie energy(products)-energy(reactants) < 0) and the reaction releases energy to the surroundings (they get hot). If the energy of the products is greater than that of the reactants, the reaction is **endothermic**, and ΔH is positive (ie energy(products)-energy(reactants) > 0) and the reaction must take energy from the surroundings (they get cold).

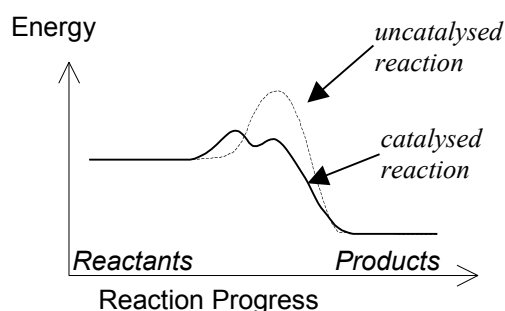
If the reaction in the forward ("left to right") direction is exothermic. The reaction in the reverse ("right to left") direction is endothermic.



Reactions that are exothermic in one direction are endothermic in the other and, in general, reactions can go both forwards and backwards, that is the products of a reaction can react together to reform the reactants, and this leads to the idea of chemical equilibrium.

Catalysts

The speed of a reaction depends on the height of the activation energy barrier, the higher the barrier, the slower the reaction. In general, however, we want reactions to proceed as fast as possible. Catalysts are substances that provide alternate reaction pathways for a reaction, with lower barriers. Catalysts are not consumed in the reaction and do not alter the final yield of products, they only make the reaction go faster by providing a new, lower energy, pathway for the reaction.



There are many different types of catalyst. For example the catalytic converter in a car is a canister in the exhaust system containing palladium oxide. The palladium oxide catalyses the conversion of carbon monoxide and unburnt petrol to water and carbon dioxide and converts nitric oxide (a component in photochemical smog) to nitrogen gas. The canisters cost about \$500 and are "poisoned" by lead - which is why you should never put leaded petrol in a car that takes unleaded fuel. Catalysts are used in virtually all chemical reactions in industry, for example making foams, plastics and polymers, lubricants, margarine, fertilisers... Chlorofluorocarbons (CFCs) catalyse the destruction of ozone in the atmosphere. Nature has also developed catalysts for most biological reactions. **Enzymes** are biological catalysts whose function has been perfected by evolution.

Acids and Bases

Acids and bases are common everyday chemicals in the home and play an indispensable role in industry and the laboratory. Some common acids and bases and their household uses are listed below.

Acid	Formula	Use	Base	Formula	Use
Acetic Acid (vinegar)	CH ₃ COOH	flavouring, preservative	Sodium Hydroxide (lye)	NaOH	oven cleaner, unblocking plumbing
Citric Acid (lemon juice)	H ₃ C ₆ H ₅ O ₇	Flavouring	Ammonia	NH ₃	household cleaner
Phosphoric Acid	H ₃ PO ₄	rust remover	Sodium Carbonate	Na ₂ CO ₃	water softener, grease remover
Hydrochloric Acid	HCl	tile cleaner	Sodium Hydrogen Carbonate	NaHCO ₃	rising agent in cakes (baking soda), mild antacid
Acetylsalicylic Acid	C ₉ H ₈ O ₄	Aspirin	Trisodium Phosphate	Na ₃ PO ₄	cleaner for surfaces before painting
Ascorbic Acid	H ₂ C ₆ H ₆ O ₆	vitamin C			
Carbonic Acid	H ₂ CO ₃	soda water			
Boric Acid	H ₃ BO ₃	mild antiseptic			

There are many ways to define acids and bases. The definitions could be phenomenological, that is, they depend on the observable physical or chemical properties of acids and bases:

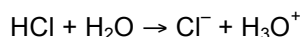
Properties of Acids

Sharp, sour taste
 React with metals to produce H₂ gas
 Corrosive
 Change litmus indicator to red

Properties of Bases

Caustic effects on skin
 Slippery feel
 Dissolve fats and grease
 Change litmus indicator to blue

We can also define acids and bases in terms of their molecular structure. Acids are substances that dissolve to produce H⁺ ions (really H₃O⁺ ions) they are **proton donors**. An acid donates a proton to water:



The hydrochloric acid (HCl) acts as an acid by donating an H⁺ to water leaving the Cl⁻ ion; the water acts as a base by accepting the H⁺ to form H₃O⁺.

Bases are substances that dissolve to produce OH⁻ ions; bases are **proton acceptors**. A base accepts a proton from water:



Ammonia (NH₃) acts as a base by accepting a proton from water to form the ammonium ion (NH₄⁺). The water acts as an acid to donate the H⁺ ion and form OH⁻. Water acts as a base in the first case and an acid in the second. Substances that can act as either a base or an acid are **amphoteric**.

Demonstrations

Demonstration/Example	Purpose
Light Sticks	To illustrate that reactions go faster when they're heated up
Chemical Hot and Cold Packs	To demonstrate exothermic and endothermic reactions (and also the fact that many chemical reactions are reversible)
Decomposition of H ₂ O ₂	To illustrate different catalysts
'talking to' a flask	To illustrate how carbon dioxide dissolves in water producing an acidic solution
Natural indicators	To illustrate how ubiquitous acids and bases are in nature

Learning Objectives

At the end of this topic you should have gained the following knowledge:

- Know that when a chemical reaction occurs, energy is required to break bonds and is released when new bonds form
- Know that almost all chemical reactions involve an energy (temperature) change
- Know that both endothermic and exothermic reactions require some energy to initiate them: the *activation energy* for the reaction
- Know that reactions can go both forwards and backwards
- Understand the role of catalysts in providing alternate reaction pathways with lower activation energies so that it is easier for a reaction to proceed
- Know the meanings of the terms *acid* and *base*
- Be able to identify some common acids and bases
- Know what an *amphoteric* substance is

Questions

1. What is meant by the terms exothermic and endothermic reactions?
2. Why is it that a reaction that is exothermic, such as burning natural gas on a gas stove, requires heat to be input (a match or a spark) before it starts?
3. Explain the role of a catalyst.
4. Why do reactions go faster as the temperature increases?
5. What do the following terms mean. Give an example of each:
 - acid
 - base
 - amphoteric

Things you might like to explore with your students:

Vinegar and Carb Soda Volcano: Combine equal parts vinegar and baking soda (also called sodium bicarbonate or carb soda) in a small plastic cup and watch the chemical reaction. You can explain that the little bubbles they see are actually carbon dioxide. Since carbon dioxide is heavier than air, it flows downward, giving it the lava effect. Consider adding food colouring or building volcanoes out of clay or dirt around it for added enjoyment. For instance you (or your class) could build a volcano around a plastic bottle, eg with paper maché or homemade playdough: 6 cups plain flour, 2 cups salt, 3 Tablespoons cooking oil, 3 Tablespoons cream of tartar (optional, it's a preservative and makes the dough smoother), 2 cups water + 2 food colouring; mix together until smooth, add more water if necessary, you want it fairly stiff. Almost fill the bottle with warm water and a bit of orange and/or red food colouring (for effect), add a little bit of detergent and about 2 Tablespoons of baking soda (bicarbonate of soda). Let the water cool (it was warm to make it easier to dissolve things – you can mix this up before moulding the playdough volcano). Slowly add vinegar... The detergent makes a foam with the bubbles of carbon dioxide so this looks very impressive! Although it's not "real" lava you could use it as an introduction to volcanos or to chemical reactions. You could investigate whether the temperature of the water mixture affects the result, whether the shape of the

volcano makes a difference, whether you can slow the “lava” down by adding something to it (eg gelatine, glycerine...) and what combination of vinegar and carb soda gives the best eruption.

Invisible Writing: Invisible inks can be heat activated, chemically activated or light activated. Heat activated inks include: any acidic fruit juice (lemon or apple is probably best, orange juice is a bit orange...), onion juice, baking soda, mixed in equal parts with water, vinegar, white wine, milk, sugar solution. Write or paint your message onto a piece of paper. You can use a paint brush, cotton bud, or finger to write your message. To reveal the message iron the paper, place it in an oven (less than 200 °C) or hold it up to a light bulb or other heat source. This will turn the writing brown. The paper turns brown because the ink weakens the paper so that it burns first where the writing is.

Inks developed by a chemical reaction are sneakier (you need to know what chemical to use!). Most of these are acid or base indicators and can be “developed” by spraying with an acid or base or indicator solution. Some of them double up as heat activated inks. They include phenolphthalein, which can be developed using ammonia fumes or spraying with carb soda solution, vinegar which can be developed by spraying with red cabbage water, ammonia, developed by red cabbage water, starch (eg the water from rinsing rice or boiling potatoes), developed by iodine solution (iodine from a pharmacy). There are more but they involve harder-to-get chemicals (or chemicals that can cause nasty stains – like salt and silver nitrate). Inks that can be seen using ultraviolet light (black light like at pubs) tend to also become visible when heated. These are dilute laundry liquid or laundry powder dissolved in water – the “bluing” agent glows, tonic water (the quinine glows) and vitamin B12 dissolved in vinegar.

Rusty Nail/Steel Wool: This demonstrates that the chemical reaction of rusting changes the structure of the nail or the steel wool (ie it becomes heavier as it rusts) and can release heat. Start with an iron nail or some steel wool and weigh it to begin with. Soak the nail/steel wool in vinegar for about a minute. This removes any protective coating and helps it start to rust. Remove from the vinegar and place the nail on some damp cotton wool (and leave to rust – this might take a while...). The steel wool is faster (why??). With the steel wool, remove it from the vinegar and wrap it around a thermometer, putting both into a sealed jar (to make the temperature change more obvious, the glass is an insulator). After a few minutes you’ll see the temperature on the thermometer start to rise (rusting is an exothermic reaction, it releases heat as new bonds are made). Once you can see a decent amount of rust, take the steel wool out of the jar, dry it and weigh it. The rusty steel wool (or nail) weighs significantly more than the original steel wool because the iron reacts with oxygen molecules in the air to form iron oxide (rust); the extra mass has come from the oxygen in the air. You can explore how long rusting takes, whether the vinegar or water was important, whether rusting is faster when it’s hot or **when it’s cold, what happens if you put it in a sealed jar, vs open to the air.**

Light Sticks: You can buy light sticks from camping stores or novelty luminescent bracelets or necklaces from \$2 shops. With 3 light sticks you can put one in ice water, another in warm water and leave the third at room temperature. Over a few minutes there will be a noticeable difference in the brightness of the light sticks (the one in warm water will glow most brightly). Try to get your students to think about what they might see before they do the experiment and then to think about explanations for what they do see. A follow up question is which light stick will glow for longest. If you do the experiment in the afternoon your students can make a prediction and then see what has happened the next morning. (The light stick in ice water will glow longest).

Instant Hot (again from camping stores or pharmacies) and Cold (from pharmacies)

Packs: Instant hot packs work by taking advantage of substances, typically salts that release a lot of heat (energy) when they dissolve in water. Hot packs typically use calcium chloride or magnesium sulfate and can reach a temperature of up to 90 °C. Cold packs use substances that absorb a lot of heat (energy) when they dissolve in water. Ammonium nitrate is the most commonly used salt. These packs can get below 0 °C. (You can use one to freeze water into ice as a demonstration – sit it on top of a little puddle of water on a piece of wood and when the water freezes you can lift the wood off the table by just holding the cold pack) In single-use hot or cold packs, the water and the salt are kept in single compartments with a thin plastic wall between them. To use the pack, the wall is broken and the salt and water are allowed to mix. Re-usable hot packs are slightly different. They consist of plastic pack containing, typically, a supersaturated salt solution and an enclosed disk. When the disk is pressed it creates a nucleating source for the salt to crystallize, eg sodium acetate will crystallize exothermically, releasing heat. The pack can be reactivated by dropping it in some boiling water or heating it in a microwave and dissolving all the salt back into a supersaturated solution. You can demonstrate a commercial hot or cold pack or dissolve some of the relevant salt in water in a container and measure the temperature with a thermometer (or touch the container!)

Natural Acid-Base Indicators: You can make a number of acid base indicators from household products:

1. Red Cabbage Indicator – boil some red cabbage in water until the water is deeply coloured, cool.
2. Red Pigment from a Rose – extract the pigment by soaking the petals in ethanol (alcohol) for an hour or so
3. Red Wine – dilute red wine approximately 1:6 with water, ie add 60 mL of water to 10 mL of wine ...
4. Make a strong cup of Black Tea and let it cool

5. Beetroot Juice, Blackberry Juice, Blueberry Juice, Cherry Juice etc (these are “reddish” for acids but “purplish” for bases)
6. Turmeric or Curry powder (the yellow pigment, curcumin, changes to reddy-brown in the presence of a base)
7. Hydrangea and Geranium petals (the petals contain anthocyanins which changes from red/orangered in an acidic solution to blue in a basic solution). Indeed lots of flowers have pigments that can act as indicators. You can play around with different flowers, extracting the anthocyanins with boiling water (they are water soluble) or ethanol (or propanol) or acetone.
8. Vanilla extract. This is an olfactory (ie smell) indicator. You can't smell the vanilla in a basic solution because the vanillin molecule ionizes although you can smell it in a neutral or acidic solution. (Onion is another olfactory indicator, you can't smell it either in basic solutions)

Using a known acid and a known base (eg use a few drops of hydrochloric acid and a few drops of sodium hydroxide, both of which you can get from a hardware store) demonstrate the colour change of each indicator that indicates acid or base. Some indicators will change colour to indicate both an acid and a base (eg the cabbage) and some will only change colour in the presence of an acid or in the presence of a base. Get the students to use the indicators to determine whether a range of household items (eg list in the notes or slides) is acidic or basic. You could also test whether soil (by mixing it with water) is acidic or basic and use this to lead into a discussion about soil degradation and land management. You could also test whether water in a swimming pool (or even out of a tap) is acidic or basic (you might be surprised)!. Another “exciting” experiment is to see whether fruit (eg a slice of apple) is acidic or basic. Drip some red cabbage juice on the apple to make it go bright pink! You can use one of the natural indicators as above or you could even make your own...

Home-Made Indicator Paper: Soak blotting paper in the red cabbage juice and then dry it. Once it's dry cut it into strips. Put the strips in a range of liquids to determine whether they are acidic or basic.

Separating the Pigments: Many of the indicators extracted from the plants above contain more than one pigment molecule. You can separate the pigments using chromatography (like in Section 1). Place a strip of coffee filter paper into the pigment solution and see whether you get multiple bands of pigment as the liquid rises into the paper.

Milk and Vinegar: Heat a cup of fresh milk to hot but not boiling and add about 2 Tablespoons of vinegar and stir for about a minute. It's amazing to witness the separation between the curds and whey. Let the milk cool down and pour it through a strainer, rinsing with water. Collect the lumpy blobs and mould them into a shape. This will harden in a few days. Adding vinegar to milk makes casein (from the latin word for “cheese”). The casein forms blobs (micelles) because it cannot mix with the water phase in the milk. You can explore whether more vinegar makes more casein, if you get the same results with low-fat milk, do all types of vinegar work, will other acids cause the same effect...

Yeast: yeast is a biological microorganism used to ferment sugar into alcohol and carbon dioxide. Put about 2 ½ cm of warm water into the bottom of a small, clean, see-through plastic bottle. Add a packet of dried yeast and gently swirl it into the water. Add about a teaspoon of sugar and swirl it around a bit more. Blow up a balloon a few times and stretch its neck a bit so you can place it over the top of the plastic bottle. Let the bottle sit in a warm place for about 20 minutes and you'll see the balloon start to inflate. As the reaction proceeds the carbon dioxide gas fills the bottle and then fills the balloon. It is this carbon dioxide that gives us the “holes” in bread. You can explore how temperature affects the yeast, how the size of the container affects the results, what combination of water and room temperature creates the most carbon dioxide, what the best “food” for the yeast is (sugar, honey, golden syrup...)

Acid Rain: This experiment uses vinegar and chalk to show how acid rain affects rock and buildings. Place a piece of chalk in a small dish (eg a saucer, to catch excess vinegar) and use an eye dropper to drop vinegar onto the chalk in a constant stream. The chalk will hiss and fizz and as carbon dioxide is produced the chalk will crumble. You could use this to initiate a discussion on pollution (acid rain) and how we can protect monuments and buildings.

Rubber Chicken Bone: a variation on the acid rain experiment is to take a clean chicken bone (leg bone), noticing how rigid it is. Put the chicken bone into a jar and immerse it in vinegar for about 3 days (put the lid on the jar, although you might need to unscrew it every so often!). After 3 days remove the bone and rinse off the vinegar. It's now very “rubbery”. Like with the chalk, the vinegar reacts with the calcium in the chicken bone, leaving only the connective tissue, collagen. Once the calcium is “dissolved” there is nothing to keep the bone hard. You could use this to introduce bones, healthy diets... You could also explore how the length of time in the vinegar affects the bone, whether smaller bones become bendier faster, whether different types of vinegar affect the bendiness... What will happen if you put the bone in bleach (hint, the bleach dissolves the collagen)?

Both these last experiments (as well as the vinegar and carb soda volcano at the top) utilise the general reaction: an acid plus a carbonate (in these latter two case calcium carbonate, and in the first experiment sodium carbonate) gives carbon dioxide gas plus a salt plus water.