

Level One Chemistry

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Chapter 1

Introduction and Preliminaries

1.1 What is Chemistry, anyway?

Chemistry is the scientific discipline involved with compounds composed of atoms, i.e. elements, and molecules, i.e. combinations of atoms: their composition, structure, properties, behavior and the changes they undergo during a reaction with other compounds. (Wikipedia)

Studying chemistry will help you understand and appreciate the world in which you live. Advances in chemistry have had an enormous influence on our modern lifestyle and standard of living. Inventions such as semiconductors, polymers, pharmaceuticals and advanced materials of all kinds are based on chemical science. The study of chemistry leads to a deep appreciation of the scientific method, particularly the intellectual skills needed to develop new theories and to design experiments to test the validity of these theories. (Chemistry Undergrad. Handbook 2018, The Univ. of Auckland)

There are five main branches of chemistry; we'll look at each in turn.

1.2 Analytical chemistry

Analytical chemistry is the study of measurements.

- How should we analyse compounds like pharmaceuticals?
- How do we know that our reaction produced what we think it did?
- How efficient is our process?

Essentially, an analytical chemist wants to find out what is in a given mixture, and how much of each component is present.

1.3 Biochemistry

Biochemistry is the branch of chemistry directly related to questions in biology.

- How do biological processes function at a chemical level?
- How is evolution shaped through chemical processes?

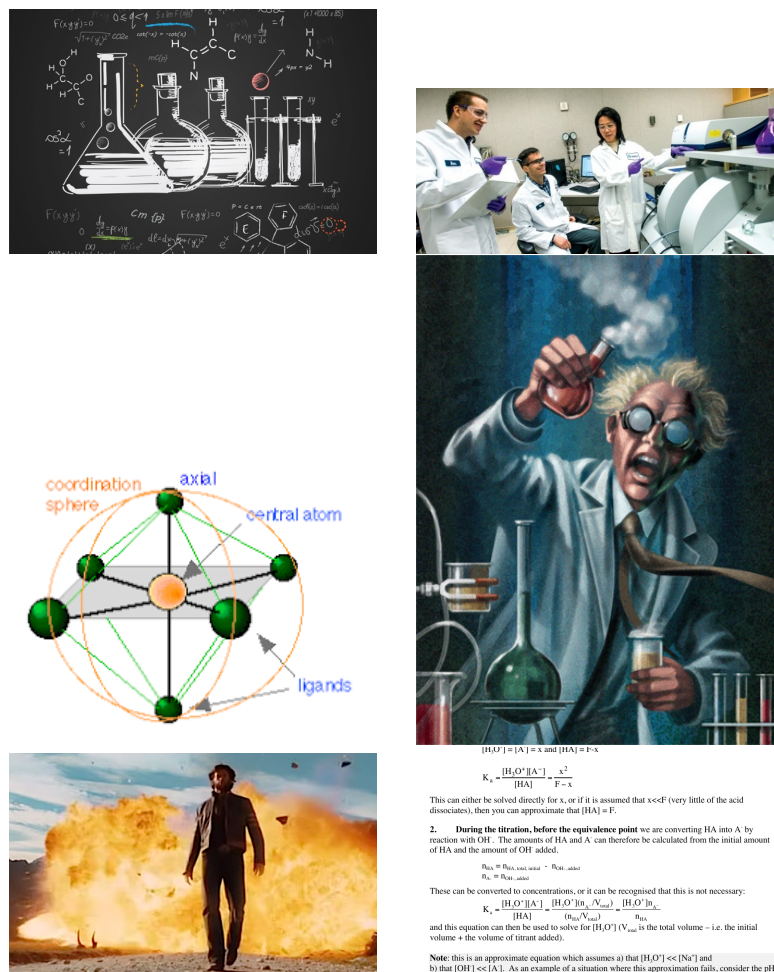


Figure 1.1: Which picture best describes chemistry?

1.4 Inorganic chemistry

Inorganic chemistry is the study of the makeup and reactions of inorganic compounds: any compound that is not composed mainly of carbon. Major areas of study include metallic chemistry, nuclear chemistry, coordination chemistry, and industrial inorganic chemistry.

- Why do the transition metals form brightly coloured solutions?
- Why do electrons sit where they do?
- What is the chemistry of uranium and other radioactive elements?

1.5 Organic chemistry

Organic chemistry is the theoretical side of biochemistry — the study of carbon-containing compounds like oils, alcohols, and fats. Organic chemists often work

with massive molecules, including polymers like plastics.

- How do we make biodegradable plastics?
- Can we make cleaner carbon-based fuels?
- How should we produce drugs efficiently?

1.6 Physical chemistry

Physical chemistry is the application of physics to chemistry. This includes the physical structure of molecules and atoms, as well as electrochemistry and spectroscopy (the study of the interaction of molecules with light).

- How can we build efficient batteries?
- How do superconductors work?

Chapter 2

Atoms and Molecules

2.1 Elements

Every substance that you can see, hear, or touch is made of atoms. To a first approximation, the subject of chemistry is simply the study of the properties of, and interactions between, atoms. Historically, the ancient Greeks postulated the existence of a smallest unit of matter but it was not until the 19th century that experimental evidence of atoms was produced.

An atom consists of a positively charged centre, called the nucleus, and a cloud of negatively charged electrons surrounding it. The nucleus is further made up of positively charged protons, and uncharged neutrons. Overall, an atom is electrically neutral: it has no net charge, because it has the same number of protons as electrons. A simple model of the atom that is a good first approximation is given as fig. 2.1.

A substance made up entirely of a single type of atom is called an element. Some examples of elements are shown as fig. 2.2.

Different atoms have different properties depending on the number of protons in their nucleus; this number is called the atomic number of the atom. The atom with atomic number 1 is hydrogen, a highly reactive gas. The atomic

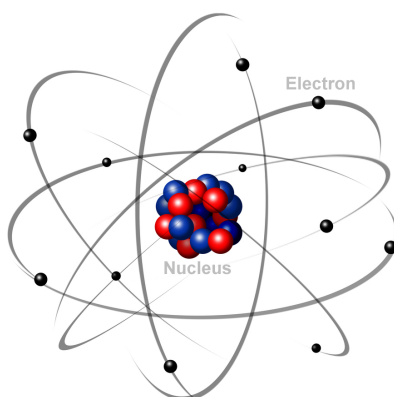


Figure 2.1: A diagram of an atom.

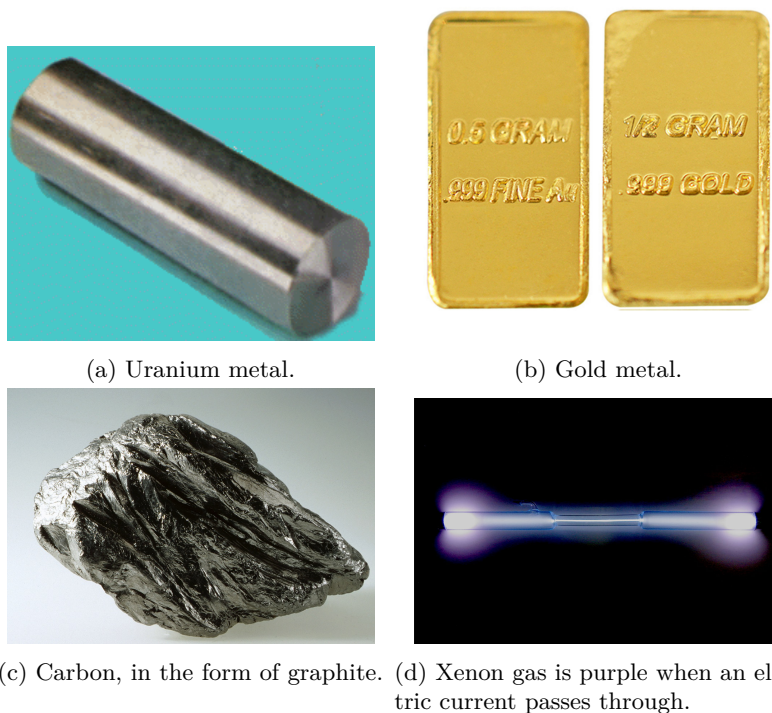


Figure 2.2: Four examples of elements.

weight of an atom is the sum of the number of protons and neutrons. Helium has two protons and two neutrons: it therefore has atomic number 2 and atomic weight 4. The type of the atom is entirely determined by the atomic number — every atom of boron has exactly five protons. However, not every atom of a particular element will always have the same number of neutrons. For example, naturally occurring uranium can have 146, 143, or 142 neutrons. Different varieties of an element with different atomic weights are called isotopes.

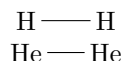
We can arrange the elements in a table, such that atoms in the same column have the same chemical properties: similar melting points, boiling points, reactivities, and so on. We obtain what is called the periodic table of the elements, fig. 2.3.¹ Each element is given a symbol made up of one or two letters; for example, manganese is Mn and iodine is I. Some elements, like gold (Au), have symbols that appear unrelated to their names: this is usually because the symbols are derived from the latin names for the elements.

- Sodium (Na): the latin name is *natrium*.
- Iron (Fe): the latin name is *ferrum*.
- Silver (Ag): the latin name is *argentum*.

¹The reason that this periodic table has non-integer atomic weights is that the weights given are an average of all the atomic weights of naturally occurring isotopes. For example, carbon has two stable isotopes, ¹³C and ¹²C.

- Tin (Sn): the latin name is *stannum*.
- Antimony (Sb): the latin name is *stibium*.
- Tungsten (W): the symbol comes from the tungsten mineral wolfram.
- Mercury (Mg): the symbol comes from the old name hydrargyrum.
- Lead (Pb): the latin name is *plumbum*.

Some elements, like hydrogen and helium, do not exist naturally as single atoms; they bond in pairs, forming a molecule.



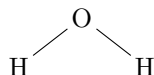
Other atoms, like carbon, form massive three-dimensional structures. Graphite, a particular arrangement of carbon, is shown in fig. 2.4.

Exercises

1. “Naturally occurring uranium can have 146, 143, or 142 neutrons.” Give the atomic weight of each isotope of naturally occurring uranium.
2. Consider the following elements: Be, Mg, Fe, Nh, In, and W. For each:
 - a) Give the elemental name.
 - b) Give the atomic number.
 - c) Rounding the atomic weights to the nearest whole number, calculate the number of neutrons and protons in the nucleus of the most common isotope.

2.2 Compounds

Of course, not everything is an element. For example, the everyday substance that we call water does not appear on the periodic table anywhere. In fact, water vapour can be formed by burning hydrogen gas in air; it turns out that water is made up of two atoms of hydrogen bonded to a single atom of oxygen. The atomic formula of water is therefore H_2O .



A substance like water that is made up purely of a particular set of atoms bonded together in a particular way is called a compound. Some other simple examples of compounds include carbon dioxide (made up of a carbon atom with two oxygen atoms), ammonia (made up of a nitrogen atom with three hydrogen atoms); an example of a larger compound is chromium(II) acetate, with atomic formula $\text{C}_8\text{H}_{16}\text{Cr}_2\text{O}_{10}$ and displayed in fig. 2.5.

One of the primary tasks of chemistry is to explain why particular compounds are seen in nature, while others are not.

Some other compounds that you should be familiar with include:

THE PERIODIC TABLE OF ELEMENTS

Atomic number

Element symbol

Element name

Relative atomic mass

Block

1

H

HYDROGEN

1.008

s

Groups of elements

s block elements

p block elements

d block elements

f block elements

Colour Key

1

H

HYDROGEN

1.008

s

1 H HYDROGEN 1.008																	2 He HELIUM 4.003	
3 Li LITHIUM 6.941	4 Be BERYLLIUM 9.012																	10 Ne NEON 20.180
5 Na SODIUM 22.990	6 Mg MAGNESIUM 24.305																	17 Cl CHLORINE 35.453
7 K POTASSIUM 39.098	8 Ca CALCIUM 40.078	9 Sc SCANDIUM 44.956	10 Ti TITANIUM 47.867	11 V VANADIUM 50.942	12 Cr CHROMIUM 51.996	13 Mn MANGANESE 54.938	14 Fe IRON 55.845	15 Co COBALT 58.933	16 Ni NICKEL 58.693	17 Cu COPPER 63.546	18 Zn ZINC 65.382	19 Ga GALLIUM 69.723	20 Ge GERMANIUM 72.640	21 As ARSENIC 74.922	22 Se SELENIUM 78.971	23 Br BROMINE 79.904	24 Kr KRYPTON 83.796	
25 Rb RUBIDIUM 85.468	26 Sr STRONTIUM 87.620	27 Y YTIUM 88.906	28 Zr ZIRCONIUM 91.224	29 Nb NIOBIUM 92.906	30 Mo MOLYBDENUM 95.940	31 Tc TECHNETIUM 97.907	32 Ru RUTHENIUM 101.072	33 Rh RHODIUM 102.906	34 Pd PALLADIUM 106.421	35 Ag SILVER 107.868	36 Cd CADMIUM 112.414	37 In INDIUM 114.818	38 Sn TIN 118.710	39 Sb ANTIMONY 121.760	40 Te TELLURIUM 127.603	41 I IODINE 126.905	42 Xe XENON 131.294	
55 Cs CAESIUM 132.905	56 Ba BARIUM 137.328	E57-71																86 Rn RADON 222.018
87 Fr FRANCIUM 223.020	88 Ra RADIUM 226.025	E89-103																118 Og OGANESSEN 284.100
57 La LANTHANUM 138.905	58 Ce CELESIUM 140.116	59 Pr PRASEODYMIUM 140.908	60 Nd NEODYMIUM 144.242	61 Pm PROMETHIUM 144.913	62 Sm SAMARIUM 150.362	63 Eu EUROPEUM 151.964	64 Gd GADOLINIUM 157.253	65 Tb TERBIUM 158.925	66 Dy DYSPROSIUM 162.500	67 Ho HOLMIUM 164.930	68 Er ERBIUM 167.259	69 Tm THULIUM 168.934	70 Yb YBBIUM 173.054	71 Lu LUTETIUM 174.967				
89 Ac ACTINIUM 227.028	90 Th THORIUM 232.038	91 Pa PROTACTINIUM 231.036	92 U URANIUM 238.029	93 Np NEPTUNIUM 237.048	94 Pu PLUTONIUM 244.041	95 Am AMERICIUM 243.061	96 Cm CURIUM 247.070	97 Bk BERKELEYUM 247.070	98 Cf CALIFORNIUM 251.080	99 Es EINSTEINIUM 252.083	100 Fm FERMIUM 257.095	101 Md MENDELIUM 258.10	102 No NIOBIUM 259.101	103 Lr LAWRENCIUM 262.103				

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Figure 2.3: The periodic table.

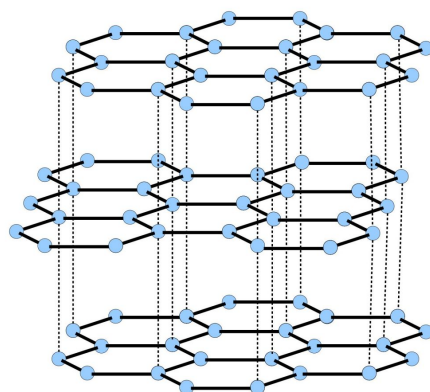


Figure 2.4: The physical structure of graphite. Blue dots represent carbon atoms.

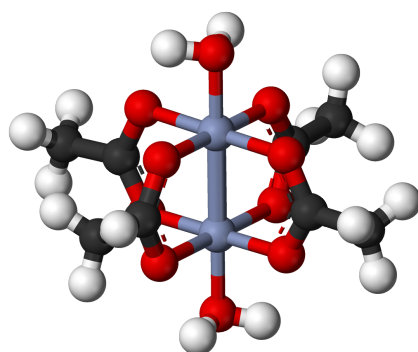
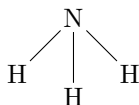


Figure 2.5: A compound of dichromium. Blue: Cr, red: O, black: C, grey: H.

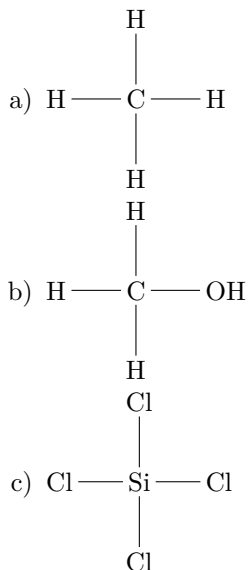
- Ammonia, NH_3 (incredibly pungent gas, used as a precursor for foods, fertilisers, and pharmaceuticals).
- Calcium carbonate, CaCO_3 (mineral found in rock).
- Sodium chloride, NaCl (table salt).
- Sodium hypochlorate, NaClO (used in bleach).
- Sodium hydroxide, NaOH (also known as lye, used as drain cleaner).
- Ethanoic acid, CH_3COOH (active ingredient in vinegar).

As you can see, there are multiple ways to write down a chemical formula. Consider ammonia; it has molecular formula NH_3 , and we can draw its structural formula as well:



Exercises

1. Give the molecular formulae for the following molecules:

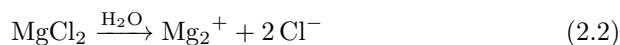


2. Ammonia is made up of nitrogen and hydrogen. Explain how ammonia gas is different from a mixture of nitrogen and hydrogen.

2.3 Ions

Atoms are always electrically neutral. However, it is possible to force an atom to gain or lose an electron; the resulting charged particle is called an ion. For example, we can remove the single electron from sodium, producing Na^+ .²

If we take a bunch of positive ions (cations) and negative ions (anions), they tend to want to clump together because opposite charges attract. A compound made up of anions and cations bonded ionically is called a salt. Some examples of salts include table salt (NaCl), hydrogen fluoride (LiF), and magnesium chloride (MgCl_2). If a salt is placed in water, it will (usually) dissociate into its component ions; for example,



You may have noticed that magnesium chloride is made up of two anions and only one cation. This is because elements in the main group of the periodic table (up to element 20, calcium) are quite well-behaved: when they form ions,

²It is even possible, though much harder, to add an electron to sodium and form Na^- . The result is quite unstable.

12: Magnesium

2,8,2

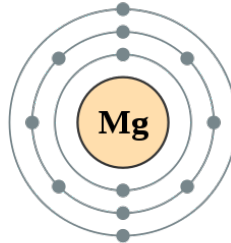


Figure 2.6: The electron shells of a magnesium atom.

20: Calcium

2,8,8,2

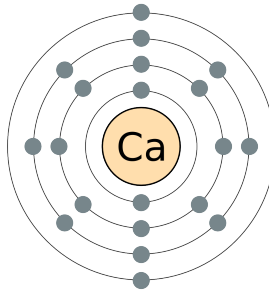


Figure 2.7: The electron shells of a calcium atom.

they gain or lose enough electrons to end up at the nearest noble gas (element in the final column, like argon or neon). Why is this? It turns out that electrons tend sit at different orbits around the nucleus; let's take magnesium as an example (fig. 2.6).

As you can see, the first two electrons sit the closest to the nucleus; then there is a second shell, which can fill up with the next eight electrons; and in magnesium the third, outermost shell (the valence shell) fills up with the remaining two electrons. We say that the electron configuration of magnesium is 2,8,2.

Another example is calcium (fig. 2.7), which has an electron configuration of 2,8,8,2.

The electron structure of atoms can get much more complicated; however,

there will always be two simple rules:

- Atoms and ions with a full outer shell are very stable.
- Atoms and ions tend to want to gain or lose electrons in order to obtain a full outer shell.

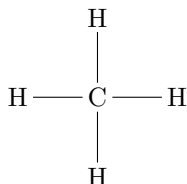
Hence magnesium and calcium tend to want to lose their two valence electrons in order to obtain a full outer shell; chlorine and fluorine tend to want to gain an electron; and noble gases are incredibly stable, and so tend not to react with anything.³

Exercises

1. Give the electron configuration of the following atoms and ions:
 - a) Na^+
 - b) B
 - c) Ar
 - d) Al^{3+}
 - e) O^{2-}
 - f) H^+
 - g) H^-
2. Explain why argon does not tend to form stable compounds.
3. Give two examples of atom-like systems with the electron configuration 2,8.
4. These chemicals are missing the correct subscripts; complete the formulae given.
 - a) MgCl_x
 - b) AlF_y (in this compound, the aluminium ion has charge 3+)
 - c) H_zCl

2.4 Covalent Bonds

Atoms in the centre of the main group, like carbon and nitrogen, tend not to bond ionically because they are ‘too far’ away from being ionic. Instead, they tend to form covalent bonds: they share electrons with other atoms. Compounds made up of covalently bonded atoms and ions are called molecules.

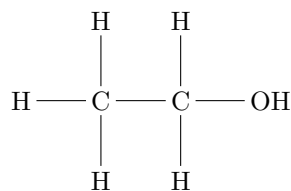


³That's not to say that they *never* react: with some trouble it is possible to synthesise XeF_4 , for example.



Figure 2.8: The formation of iron(III) thiocyanate from Fe_3^+ ions in solution.

Methane (CH_4) is an example of a covalently bonded molecule; it is also an example of an organic molecule (one which contains carbon). Each bond consists of two electrons, shared between the carbon and hydrogen atoms.



Ammonia, chlorine gas (Cl_2), oxygen gas, and ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) are also examples of covalently bonded molecules.

Many covalent compounds are also ions. For example, the nitrate ion NO_3^- consists of covalently bonded atoms; some others include:

- Sulfate, SO_4^{2-}
- Hypochlorate, ClO^-
- Carbonate, CO_3^{2-}
- Hydrogen carbonate, HCO_3^-
- Iron(III) thiocyanate, $[\text{FeSCN}(\text{H}_2\text{O})_5]^{2+}$ (??)

Exercises

1. Using your chemical knowledge, classify the following as ionic or covalent; if a substance is ionic, give the cation and anion.
 - a) Ethanol, $\text{CH}_3\text{CH}_2\text{OH}$
 - b) Sodium bromide, NaBr
 - c) Calcium carbonate
 - d) Methane, CH_4
 - e) Aluminium oxide

2.5 Chemical reactions

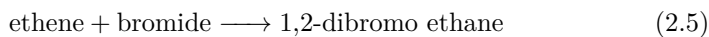
The most interesting properties of elements and compounds come from their reactivities. A chemical reaction is a process which takes a set of particles (the reactants) and rearranges the atoms and bonds to form a new set of particles (the products). We often summarise a given reaction by writing a chemical equation:



For example, consider the following:

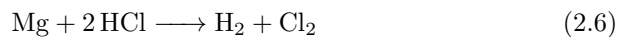


The corresponding word equation is:



From these equations, we see that in this reaction two molecules exchange bonds to form a single new molecule.

A chemical reaction can never create or destroy atoms, so the same number of each atom must occur on each side. In the following reaction, multiple of the same molecule are needed to balance the equation:



Chapter 3

Acids and Bases

3.1 Acid-base reactions

We need the following definition:

An **acid**¹ is a chemical species² that tends to donate protons when in solution.

Recall that a proton is exactly the same thing as an H^+ ion. An example of an acid is hydrogen chloride (usually known as hydrochloric acid), HCl . In water, it dissociates into ions:



In other words, the HCl donates its proton into solution in the form of an H^+ ion.³

Ammonia (NH_3), a colourless gas, also acts as an acid due to the following reaction:



A dual definition can also be made.

A **base**⁴ (or an **alkaline species**) is a chemical species that tends to accept protons when in solution.

The main example of a base is the hydroxide ion OH^- ; if it meets a proton swimming around in solution, it will combine with it to form a water molecule:



In reality, most of us do not keep a stock of hydroxide ions sitting around by themselves. Most bases that we use are salts that dissociate in water to produce hydroxide ions, which then accept protons. Examples of bases include

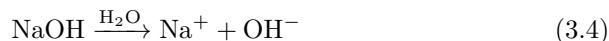
¹More specifically, a Brønsted-Lowry acid.

²A chemical species is any thing that exists on its own in solution — for example, an ion or a molecule.

³We're actually lying to you here: there is no such thing as an H^+ ion in water. In fact, the HCl donates its proton to a water molecule which acts as a base, forming an hydronium ion ($\text{HCl} + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$).

⁴More specifically, a Brønsted-Lowry base.

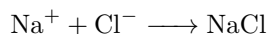
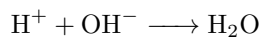
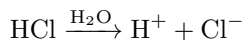
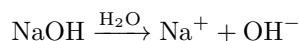
NaOH and KOH:



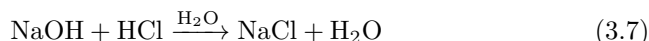
Ammonia can also act as a base:⁵



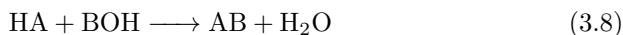
If we place sodium hydroxide and hydrochloric acid together in solution, the following set of reactions will occur:



So in total, the following is observed:



By the same reasoning, the following rule can be easily justified (where HA is some acid, and BOH is some base):



In other words, acid plus base forms salt plus water.

Exercises

- Complete the following reaction schemes, making sure to balance where necessary:
 - $\text{H}_2\text{SO}_4 + \text{NaOH} \longrightarrow$
 - $\text{HNO}_3 + \text{CH}_3\text{NH}_2 \longrightarrow$
 - $\text{HCl} + \text{KOH} \longrightarrow$
- Describe the reaction of ammonia with KOH.

3.2 Measuring acidity

Important concept: the strength of an acid is measured by the concentration of protons that it donates.

We measure the ‘acidity’ of a solution using a scale known as the pH scale — more acidic substances have a lower pH value. The acidity of a substance can be measured by adding universal indicator and noting the colour change (fig. 3.1), by using litmus paper (acid turns blue litmus paper red, and base turns red litmus paper blue), or by using an electronic pH meter.

⁵Water is another compound which can act as both an acid and a base.



Figure 3.1: The colours of universal indicator paper.

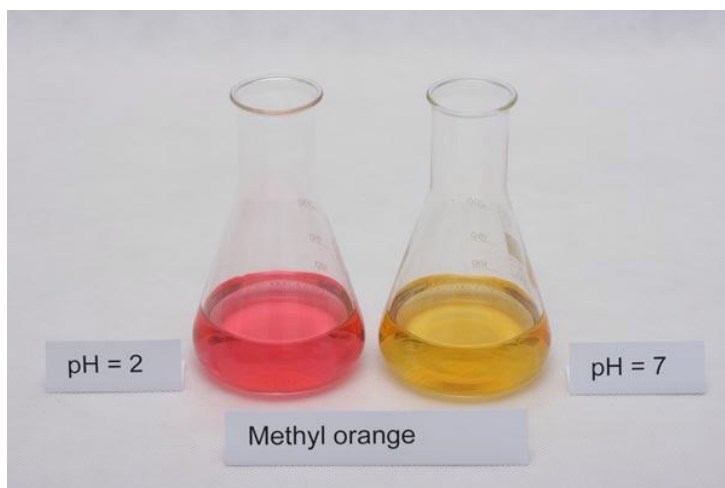


Figure 3.2: Another indicator is methyl orange; there are many others, and the choice of indicator depends on the particular reaction that is being carried out.

pH value	Solution	Universal indicator colour
1	(<i>acidic</i>) HCl	red
2	Lemon juice	
3	Ethanoic acid	
6	Milk, urine	green
7	(<i>neutral</i>) Pure water, blood	
8	Sea water	
9	Baking soda (NaHCO ₃)	purple
11	Ammonia	
13	NaOH	
14	(<i>basic</i>) H ₂ SO ₄	

While most solutions you meet will have a pH in the range 1-14, it is important to realise that it is entirely possible to have a solution with a pH outside that range; highly concentrated acids can have a negative pH value, for example.

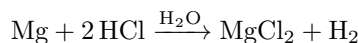
Exercises

1. Chlorine water is formed by adding chlorine gas to water. Some of the chlorine dissolves in water and some reacts with water. Explain what happens to a piece of blue litmus paper when it is used to test the chlorine water.
2. HCl is sitting in aqueous solution. A small amount of universal indicator is added, and then sodium hydroxide is added dropwise. Describe and explain any observations.

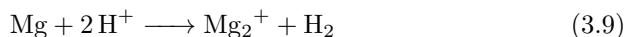
3.3 Acid-metal reactions

A **metal** is an element that is shiny and generally hard when a solid, and is malleable (easy to bend), ductile (can be drawn into wires), and a good electrical conductor.⁶

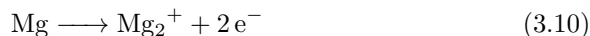
If we drop a piece of metal into acid, then we observe bubbling on the surface of the metal. This is because the metal is donating electrons to the acid, allowing the hydrogen cations to form covalent bonds and escape as a gas. Let's look at an example:



Getting rid of the Cl⁻ anions that don't actually take part in the reaction, we have the following reaction:



We can split this up further, by looking at where the individual electrons go. There are two reactions happening at the same time:



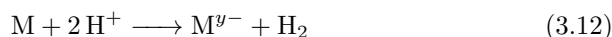
⁶This is a very naive definition; more correctly, a metal is an element that has free valence electrons.



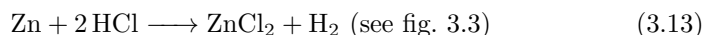
Figure 3.3: The reaction of zinc metal with hydrochloric acid.

This kind of reaction is an oxidation-reduction reaction (or redox reaction). This year, we only care about the specifics of acid/metal reactions, but more interesting redox reactions can occur; for example, the underlying mechanism of a battery is a redox reaction.

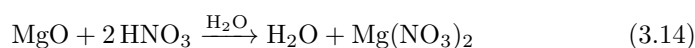
In general, we have the unbalanced equation (where M is a metal, and y is the absolute charge on the cation of M that is formed)



In other words, metal plus proton gives metal ion plus hydrogen (or metal plus base gives salt plus hydrogen). One final balanced example:



Note that usually energy is released from a metal/acid reaction as heat. Acids also react with metal oxides. For example,



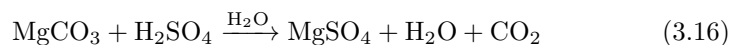
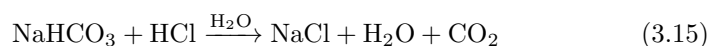
In general, a metal oxide plus an acid produces a salt and water.

Exercises

1. Small pieces of zinc were dropped into a test tube of dilute sulfuric acid solution. Describe the observations you would make of this reaction.
2. Magnesium is added to nitric acid. Write down a balanced equation for this reaction.

3.4 Acid-carbonate reactions

Carbonates and hydrogen carbonates also react with acids. Like with metals, a gas is produced; however, this time it is carbon dioxide (CO_2) that is produced. Two examples:



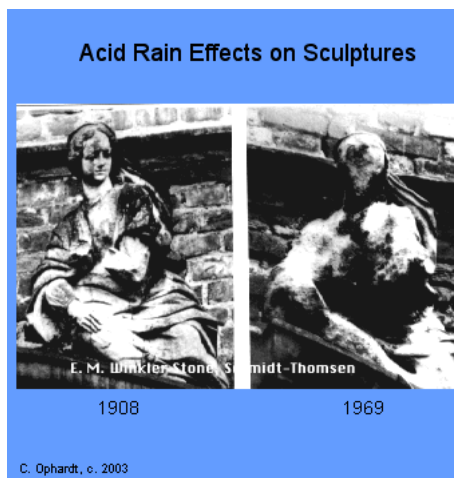


Figure 3.4: The effect of acid rain on a sculpture.

Again, this is an example of a redox reaction.

Exercises

1. Limestone is primarily made up of calcium carbonate. When acid rain (containing a component of sulfuric acid) falls on sculptures or buildings built of limestone, what would you expect to observe over time? (fig. 3.4)
2. Explain why eggshells soften in vinegar.

Chapter 4

Solubility and States of Matter

4.1 Solutions

A solution is a mixture of two substances, such that one substance (the solute) is dissolved in the other (the solvent). Common solvents include water and ethanol.

Most salts are highly soluble (dissolve easily) in water, because water is polar — water molecules have one end with a slightly negative charge and one end with a slightly positive charge, so positive ions and negative ions are attracted to different ends of the water molecules and so are split up. A diagram showing this process for NaCl is given as fig. 4.1; the negative ends of the water molecules are marked by δ^- , and the positive ends by δ^+ .

At room temperature, we have the following solubility rules for salts:

Nitrates (NO_3^-)	All soluble
Chlorides (Cl^-)	All soluble except AgCl and PbCl_2
Iodides (I^-)	All soluble except AgI and PbI_2
Sulphates (SO_4^{2-})	All soluble except BaSO_4 , PbSO_4 , and CaSO_4
Hydroxides (OH^-)	All insoluble except KOH and NaOH
Carbonates (CO_3^{2-})	All insoluble except K_2CO_3 and Na_2CO_3

It is important to realise that all compounds are soluble to some degree; calcium carbonate (limestone, CaCO_3) is insoluble in water according to the rules above, but this means that very little compound will dissolve — not none! Conversely, even if a compound is highly soluble there is a maximum amount

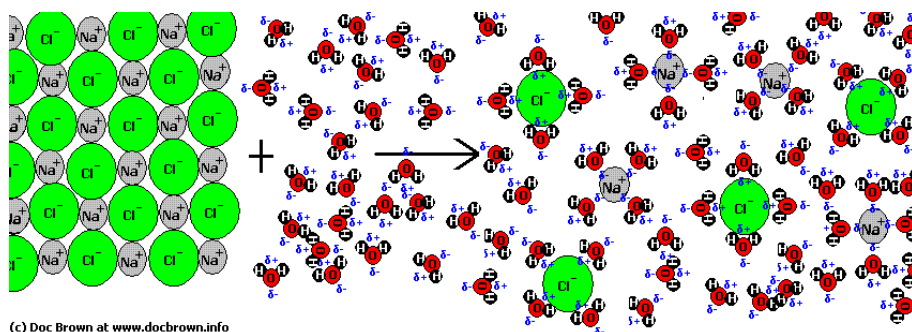


Figure 4.1: The dissolution of table salt in water.

that can be held in solution: if one continues to add salt to water, eventually it will cease to dissolve. A solution containing the maximum amount of solute for the given temperature and amount of solvent is called saturated.

The solubility of a solute can be increased by:

- Increasing the temperature of the solvent. (Warm water dissolves more solvent.)
- Reducing the common-ion effect (e.g. if NaOH is added to a solution that already contains OH^- ions, then it has reduced solubility; conversely, removing OH^- ions will increase the solubility of Na^+ ions from the NaOH).
- Increasing pressure on the solution (if we are trying to dissolve a gas like HCl).

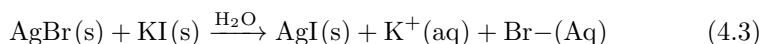
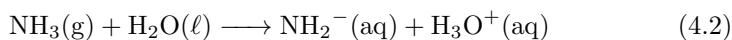
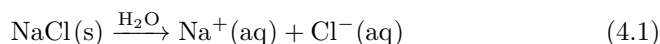
Exercises

1. Classify as soluble or insoluble in water:
 - a) KI
 - b) KNO_3
 - c) $\text{Pb}(\text{OH})_2$
 - d) Ag_2SO_4
 - e) CaCO_3

4.2 States of Matter

Obviously the behaviour of substances when we try to dissolve them depends on the state which they are in: solid, liquid, or gas (fig. 4.2). When we write chemical equations, we place the state of each product and reactant in brackets behind them: (s) — solid; (l) — liquid; (g) — gas; (aq) — aqueous (in solution with water, like fig. 4.1).¹

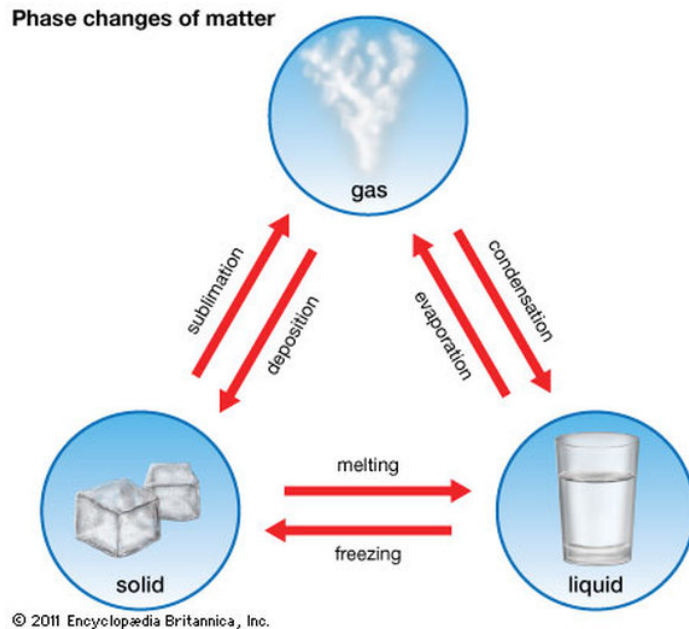
Let's look at a few examples of equations:



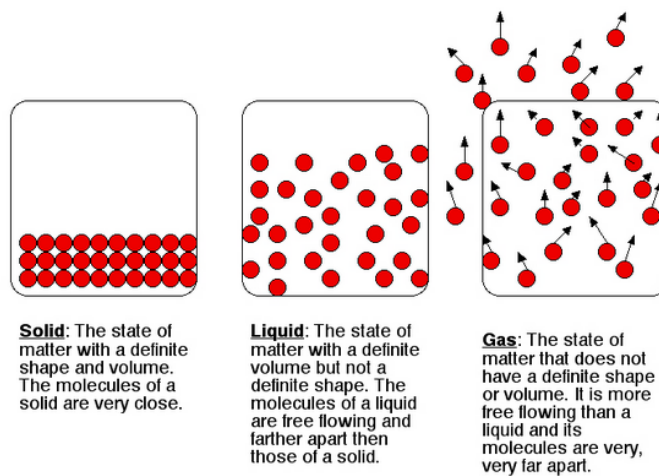
In order to change matter from one state to another, we must either put energy in or take energy out. A process which requires energy to be input to occur is called endothermic; a process which outputs energy is called exothermic. Melting and evaporation are endothermic processes, because we need to input energy into the system to make them happen; condensation and freezing are exothermic processes, because energy is released by the system when they occur.

There are, in fact, four fundamental states of matter: apart from solid, liquid, and gas, one can produce plasma (a state in which the electrons are

¹One may also see (alc), for substances dissolved in ethanol or other alcohols.



(a) Transitions between states.



(b) Physical properties of states.

Figure 4.2: The three common states of matter.

1 H																	2 He																									
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne																									
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar																									
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr																									
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe																									
55 Cs	56 Ba											72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn																
87 Fr	88 Ra											104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo																
																		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu										
																		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr										

Figure 4.3: The periodic table by state (green: solid, yellow: gas, red: liquid).



(a) Liquid mercury.



(b) Liquid bromine.

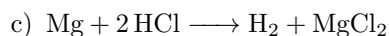
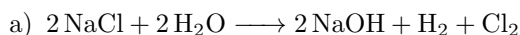
Figure 4.4: The only two elements which are liquid at room temperature.

ripped entirely free from the nucleus) by heating gas. Plasma is generated when lightning strikes, and within neon lights. Nuclear fusion normally takes place in a plasma of ^2H and ^3H , like that found in the sun.

We can classify the elements by their state at room temperature (fig. 4.3). Most elements are solid; the noble gases and some other elements in the upper right portion of the table are gaseous at room temperature; and only two elements (mercury and bromine, fig. 4.4) are liquid.

Exercises

1. Add the states of matter to the following equations:



2. Classify the following as exothermic or endothermic:

- a) Sublimation.
- b) Deposition.
- c) Cooking an egg.
- d) Reacting an acid with a metal.

Chapter 5

Rates of Reaction

5.1 Collision theory

One model of chemical reactions is collision theory. According to this model, reactants can combine to form products if:

- The reactant particles collide with enough energy (the activation energy); and
- The reactant particles collide in the correct orientation.

Graphical depictions of these requirements for a successful collision and therefore reaction can be seen in fig. 5.1.

It immediately follows that in order to increase the rate of a reaction, we can take the following steps:

- Increase the concentration of the reactants and/or increase the pressure of any gaseous reactants (so that the frequency of particle collisions is increased, and hence more successful collisions occur per second).
- Increase the temperature of the reactants (so that the speed of particles is increased, and more collisions occur with enough energy).

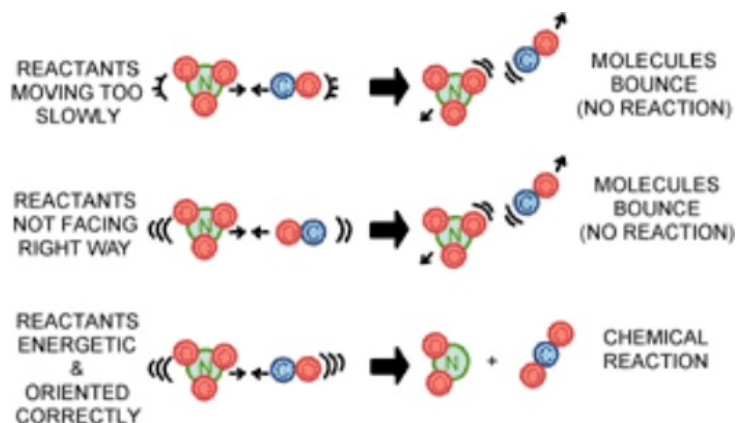


Figure 5.1: Collision theory.

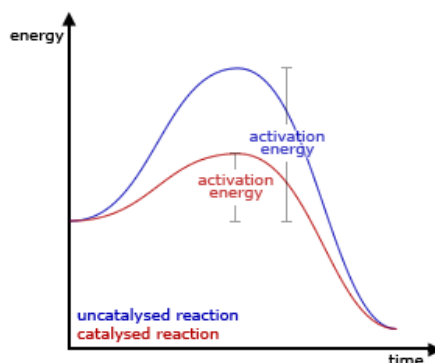


Figure 5.2: A reaction profile with and without a catalyst.

- Add a catalyst to the reaction (which provides an alternate reaction pathway with lower activation energy).
- Increase the surface area of any reactants in solid form (so that more reactant particles are exposed, and collisions can occur more frequently).

5.2 Catalysts and reaction profiles

A reaction is exothermic if the total amount of energy of the products is less than the total amount of energy of the reactants. This energy is stored in the bonds between the atoms in the compounds, and is known as enthalpy.

In order for a reaction to occur, some energy must be provided to ‘get it going’ — this is the activation energy mentioned above. The two diagrams in fig. 5.2 show the enthalpy stored in a given system during a reaction; since the energy of the reactants is less than the energy of the products, the reaction is exothermic.

The diagram also shows the action of a catalyst. Reactions usually occur in steps, producing intermediate products (like carbocations) by moving electrons around; a catalyst simply enables a different pathway of intermediate steps that requires a lower activation energy than the non-catalysed reaction path. Catalysts are not used up in reactions.

Some examples of catalysts:

- Platinum is used in car exhausts as a catalyst to break down some nasty products of combustion: $2\text{CO} + 2\text{NO} \xrightarrow{\text{Pt}} 2\text{CO}_2 + \text{N}_2$.
- Chlorine free radicals (Cl^\bullet , a chlorine atom that has lost a particular electron) are formed by the action of UV light on chlorofluorocarbons, and catalyse the breakdown of ozone through the following pair of reactions: $\text{Cl}^\bullet + \text{O}_3 \longrightarrow \text{ClO}^\bullet + \text{O}_2$ and $\text{ClO}^\bullet + \text{O} \longrightarrow \text{Cl}^\bullet + \text{O}_2$ (so the observed reaction is $\text{O}_3 + \text{O} \xrightarrow{\text{Cl}^\bullet} 2\text{O}_2$).
- Polypropylene is produced using a titanium-based catalyst.