

SCIENCE COURSE

Chemistry Fundamentals

Matter, atoms, bonds and reactions

10 Chapters

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Generated with Claude

CHEMISTRY FUNDAMENTALS

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Chemistry Fundamentals

Course 1 · Chapter 1 · Matter and Measurement: What Everything Is Made Of

According to a tale told by the Roman writer Vitruvius, the king of Syracuse suspected that a goldsmith had cheated him by replacing some of the gold in a votive wreath with a cheaper metal. He asked Archimedes to find out without damaging the wreath. Archimedes is said to have noticed, while lowering himself into a bath, that the water level rose, realised that the displaced water could give the wreath's volume, and run through the streets shouting "Eureka!". Wikipedia's Density article calls the story "well-known but probably apocryphal", and says it first appears in Vitruvius about two centuries later. Whether or not it happened, the idea in it is the starting point of chemistry: different materials have measurable properties, and you can tell them apart by measuring them. This chapter sets out what chemistry studies and the vocabulary the rest of the course uses.

What Chemistry Is

Wikipedia's Chemistry article opens with a short definition: "Chemistry is the scientific study of the properties and behavior of matter." It adds that chemistry is a physical science covering the composition and structure of matter and the changes matter undergoes in reactions, and that it addresses the nature of chemical bonds. It lists five main branches: analytical, biochemistry, organic, inorganic and physical chemistry (the picture caption adds applied chemistry as a sixth). This course is about the ideas that all of them share.

Matter and Its States

Matter is "anything that has mass and occupies space by having volume", in the wording of Wikipedia's Matter article, which also says physics has no single agreed definition and that the answer depends on the scale and the field. For chemistry the everyday version is enough. The same article describes the "classical everyday phases such as solid, liquid, and gas", and lists others: plasma, Bose-Einstein condensates, fermionic condensates and quark-gluon plasma.

Solid

Keeps its own shape and volume. The particles are packed closely and mostly stay

Liquid

Keeps its volume but takes the shape of its container. The particles are close but can

in place.

move past each other.

Gas

Fills whatever container it is in. The particles are far apart and move freely.

The descriptions of the particles in the three cards are the standard textbook picture; I have written them from general knowledge, and Chapter 2 supplies the particle model they rest on (check against your own textbook). A change between states, such as ice melting or water boiling, rearranges the particles but does not make a different substance.

Substances and Mixtures

Wikipedia's Chemical substance article defines a substance as "a unique form of matter with constant chemical composition and characteristic properties". Compare:

Term	The article's wording	Example
Element	"a chemical substance made up of a particular kind of atom"	Iron, oxygen gas
Compound	"a chemical substance that is composed of a particular set of atoms or ions"	Water, table salt
Mixture	"more than one chemical substance" with no "fixed composition"	Air, sea water, soil

The examples in the right-hand column are mine, chosen as standard cases. The article states the law of constant composition as "All samples of a compound have the same composition", and says there are 118 known elements, about 80 of them stable. It notes that a mixture's parts are combined physically, with no chemical reaction, so they can be separated by "mechanical processes, such as chromatography, distillation, or evaporation"; its example of a magnet pulling iron away from sulfur shows the idea. When a mixture is separated so that one substance is isolated, the Matter article says the result is "chemically pure".

Why "pure" is a chemist's word

> In everyday speech, "pure orange juice" means nothing has been added. In chemistry, a pure substance has one composition throughout, and juice, which contains water, sugars, acids and many other things, is a mixture however it is labelled.

Physical and Chemical Change

Melting ice and rusting iron are both changes, but they differ. In a **physical change** the same substance is still there afterwards, perhaps in a different state or shape. In a **chemical change** (a chemical reaction) one or more new substances form. The Chemistry article describes reactions as transformations that rearrange electrons in bonds to form different substances; it does not state the physical-versus-chemical distinction in so many words, so the two-way split above is the usual textbook one, given here from general knowledge. Clues that a chemical change may have happened include a colour change, a gas given off, a solid forming in a liquid, or heat or light released, though none of them is proof on its own (check; a boiling kettle gives off a gas, and no new substance forms).

Measuring Matter: Mass, Volume and Density

Two measurements describe how much matter there is: **mass** (how much matter) and **volume** (how much space it takes). Neither identifies a material, since a kilogram of lead and a kilogram of feathers have the same mass. Their ratio does. The Density article defines density as "mass per unit volume", written $\rho = m / V$, with the SI unit kilogram per cubic metre.

Material	Density (kg/m^3), from the Density article
Ice (below 0 °C)	916.7
Water at 20 °C	998.2071
Water at 4 °C	999.9720
Aluminium	2,700
Iron	7,870
Lead	11,340

Worked example: why ice floats

Ice is less dense than liquid water (916.7 against 999.972 kg/m³ at 4 °C), so it floats. Dividing, 916.7 / 999.972 = 0.9167, which means about 92 per cent of a floating block of ice is under the surface. That last step uses the idea, standard in physics, that a floating object sinks until it displaces its own weight of water; the density article's table does not state it (check). I did the arithmetic with a calculator.

Watch the units

A density of 2.7 g/cm³ is the same as 2,700 kg/m³, because 1 g/cm³ equals 1,000 kg/m³. Mixing the two forms is a common source of answers that are wrong by a factor of a thousand. Write the unit next to every number and convert before you compare.

Where to Read More

- **Encyclopaedias.** Wikipedia's Chemistry, Matter, Chemical substance and Density articles are the sources used in this chapter. I read the Chemistry and Matter articles only partly (each to the first 100,000 characters of about 200,000), the Chemical substance and Density articles in full.
- **A school textbook.** A GCSE or equivalent chemistry textbook covers this material with diagrams and practice questions.
- **Practical.** Weighing a measured volume of water, then of cooking oil, shows density directly with a kitchen scale and a measuring jug.

Exercises

EXERCISE 1

A block of metal has a mass of 540 g and a volume of 200 cm³. Calculate its density in g/cm³ and in kg/m³, and say which metal in the table it matches.

→ Solution

EXERCISE 2

Classify each as an element, a compound or a mixture: (a) sea water; (b) oxygen gas; (c) table salt; (d) air; (e) iron filings stirred into sulfur powder. Give a reason for each.

→ Solution

EXERCISE 3

You have a mixture of sand, salt and iron filings. Plan a way to separate all three, using only the methods named in this chapter and ordinary kitchen equipment, and say why each step works.

→ Solution

Quick Reference

- **Chemistry:** the scientific study of the properties and behaviour of matter
- **Matter:** anything with mass that occupies space; solid, liquid, gas and others
- **Substance:** constant composition and characteristic properties; element or compound
- **Mixture:** more than one substance, no fixed composition, separable by physical means
- **Physical change:** same substance; **chemical change:** new substance
- **Density:** $\rho = m / V$; 1 g/cm³ = 1,000 kg/m³
- **Check:** the particle descriptions of the states, and the list of signs of a chemical change

Coming next

Chapter 2 looks inside matter: the atom, its protons, neutrons and electrons, and what makes one element different from another.

Chemistry Fundamentals

Course 1 · Chapter 2 · Atoms and Elements: What Makes One Element Different From Another

In the early 1900s, Hans Geiger and Ernest Marsden, working under Ernest Rutherford at the University of Manchester, fired a beam of alpha particles at a very thin metal foil, usually gold. According to Wikipedia's article on the Geiger-Marsden experiments, most of the particles passed straight through, but a small fraction were deflected sharply, some "in all directions, sometimes more than 90°". Rutherford concluded that the atom's positive charge and most of its mass sit in a tiny nucleus, with the electrons spread around it. That picture, a small dense centre and a lot of empty space, is the one this chapter fills in. It explains what an atom is built from, why the number of protons decides which element you have, and what an isotope is.

The Atom and Its Three Particles

Wikipedia's Atom article opens: "Atoms are the basic particles of the chemical elements and the fundamental building blocks of matter." It describes an atom as a nucleus "of protons and generally neutrons", surrounded by electrons. Chapter 1 called an element "a chemical substance made up of a particular kind of atom"; this chapter says what "kind" means.

Proton

In the nucleus. Charge +1 e. Mass 1.0072764665789 Da (about 1.6726×10^{-27} kg), "approximately 1836 times the mass of an electron" (Proton article).

Neutron

In the nucleus. Charge 0. Mass 1.00866491606 Da (Neutron article), slightly more than a proton.

Electron

Outside the nucleus. Charge $-1.602176634 \times 10^{-19}$ C, the same size as a proton's charge but negative. Mass about 9.109×10^{-31} kg (Electron article).

The Daltons (Da) in the cards are units of mass for single atoms and particles, where one dalton is roughly the mass of one proton or neutron. Two things follow from the numbers. First, the

nucleus holds nearly all of an atom's mass, because the electrons are about 1836 times lighter than a proton. Second, protons and neutrons are so close in mass that, for counting purposes, each can be treated as "1".

The Atom article says the total number of protons and neutrons in an atom, which it calls "nucleons", is "the mass number". In an uncharged atom, the Atomic number article says, the number of electrons equals the number of protons, so the positive and negative charges cancel. Atoms that have gained or lost electrons are called ions and are the subject of Chapter 4 (this forward pointer is mine, from the course outline).

Atomic Number: What Defines an Element

The Atom article says: "The chemical elements are distinguished from each other by the number of protons that are in their atoms." That number is the **atomic number**, symbol Z. The Atomic number article puts it this way: it "can be used to uniquely identify ordinary chemical elements", and it is "the number of protons found in the nucleus of every atom of that element".

So an atom with 6 protons is carbon, whatever else it contains, and an atom with 17 protons is chlorine. Change the number of protons and you have a different element; change the number of electrons or neutrons and you do not.

> **Mass number and the neutron count**
>
> The Atomic number article states that the atomic number Z plus the neutron number N gives the mass number A, so $A = Z + N$. Rearranged (my rearrangement, simple algebra), the number of neutrons is $N = A - Z$.

Isotopes

Because the proton count fixes the element but the neutron count is free to vary, one element can have atoms of several kinds. The Atom article says: "Atoms with the same number of protons but a different number of neutrons are called isotopes of the same element." The Isotope article names an isotope by writing the element name, a hyphen and the mass number, for example carbon-12 and carbon-14, and in symbols as ^{12}C and ^{14}C.

Isotope	Protons (Z)	Mass number (A)	Neutrons (A - Z)
Carbon-12	6	12	6
Carbon-14	6	14	8
Chlorine-35	17	35	18
Chlorine-37	17	37	20

The names and mass numbers come from the Isotope article. The proton counts (6 and 17) are the atomic numbers of carbon and chlorine, which I took from general knowledge and the periodic table rather than from the pages I read (check); the neutron column is my own subtraction, A - Z, done with Python.

The Isotope article says isotopes have the same atomic number but different mass numbers "due to different numbers of neutrons in their nuclei", and that "all isotopes of a given element have virtually the same chemical properties, they have different atomic masses and physical properties". The article identifies carbon-14 as radioactive and mentions radiocarbon dating; it gives no half-life in the part I read, so none is given here.

> **Isotope is not a different element**
>
> Carbon-12 and carbon-14 are both carbon. Different neutron counts make different isotopes of one element, and only a different proton count makes a different element.

Relative Atomic Mass: A Weighted Average

A sample of an element usually contains a mixture of isotopes, so the mass quoted for the element is an average. Wikipedia's Relative atomic mass article defines it as "the ratio of the average mass of atoms of a chemical element in a given sample to the atomic mass constant", and says that for a given sample it is the "weighted arithmetic mean" of the masses of the individual atoms, all isotopes included. "Weighted" means each isotope counts in proportion to how common it is.

The Isotope article gives chlorine as an example: a sample "contains 75.8% chlorine-35 and 24.2% chlorine-37, giving an average atomic mass of 35.5 daltons". The Relative atomic mass

article has no chlorine example, but it works through silicon, which has three isotopes:

Isotope	Mass (Da)	Abundance (fraction)	Mass × fraction (my arithmetic)
Silicon-28	27.97693	0.922297	25.80304
Silicon-29	28.97649	0.046832	1.35703
Silicon-30	29.97377	0.030872	0.92535
Total		1.000001	28.0854

Worked example: silicon

The article's own calculation reads $A_r(Si) = (27.97693 \times 0.922297) + (28.97649 \times 0.046832) + (29.97377 \times 0.030872) = 28.0854$, and it gives the IUPAC standard value as 28.0855(3). I re-ran the sum in Python and got 28.0854, matching. The three products, rounded to five decimals, are in the table. The fractions add to 1.000001 because of rounding in the published abundances (my addition of the three printed fractions). Notice that 28.0854 lies close to the heaviest-weighted isotope, silicon-28, as it should for an average dominated by one value.

The Relative atomic mass article says the quantity is a ratio, so it has no unit; the Isotope article quotes the chlorine figure in daltons. For school purposes the two are used interchangeably, which I state from general knowledge (check).

Linking Back: Atoms and the States of Matter

Chapter 1 described solids, liquids and gases by how their particles are arranged, and flagged that picture as general knowledge. The particles in question are atoms, or groups of atoms. A change of state, such as ice melting, rearranges them without changing what they are, and that is why it is a physical change. A chemical change, by contrast, rearranges the electrons that hold atoms together (the Chemistry article describes reactions this way) and produces new substances. The Atom article's States section, which might have connected atoms to states of matter directly, fell in the part of the page I did not read (the page was cut off), so this link is the course's own framing (check).

A Short History Note

The Geiger-Marsden experiments, described at the start of the chapter, were run between 1906 and 1913 according to the article. In 1909, Geiger and Marsden reported that some alpha particles scattered by more than 90 degrees, which the earlier "plum pudding" model by Thomson could not explain. Only a compact central charge could cause the rare large deflections. The article says nothing, in the part I read, about the later discovery that the nucleus contains neutrons, so this chapter does not date that step.

Where to Read More

- **Encyclopaedias.** Wikipedia's Atom, Isotope, Relative atomic mass, Atomic number, Proton, Neutron, Electron and Geiger-Marsden experiments articles are the sources used in this chapter. I read only the first 100,000 characters of Atom (of about 252,000), Isotope (of about 148,000), Proton (of about 151,000), Neutron (of about 275,000), Electron (of about 272,000) and Geiger-Marsden experiments (of about 163,000). I read the Relative atomic mass and Atomic number articles in full.
- **A school textbook.** The periodic table in any chemistry textbook lists atomic numbers and relative atomic masses for every element.
- **Practical.** Pick any element from a periodic table and work out its proton, neutron and electron counts for one isotope, using the rules above.

Exercises

EXERCISE 1

An atom of chlorine-37 has an atomic number of 17 and is uncharged. How many protons, neutrons and electrons does it have? Then do the same for an uncharged atom of carbon-14, which has atomic number 6.

→ Solution

EXERCISE 2

A sample of chlorine is 75.8% chlorine-35 and 24.2% chlorine-37. Use the mass numbers 35 and 37 as the isotope masses to calculate the weighted average, and compare your answer with the 35.5 daltons quoted in this chapter.

→ Solution

EXERCISE 3

A friend says: "Carbon-14 has two more neutrons than carbon-12, so it must be a different element." Explain, using this chapter, whether your friend is right and what does and does not differ between the two.

→ Solution

Quick Reference

- **Atom:** a nucleus of protons and (generally) neutrons, surrounded by electrons
- **Atomic number Z:** number of protons; identifies the element; equals the electrons in a neutral atom
- **Mass number A:** protons plus neutrons; A = Z + N, so N = A - Z
- **Isotopes:** same Z, different N; virtually the same chemistry
- **Relative atomic mass:** weighted average over the isotopes in a sample
- **Particles:** proton +1 e; neutron 0; electron -1 e, about 1/1836 of a proton's mass
- **Check:** proton counts for carbon and chlorine; the link to Chapter 1's states of matter; treating relative atomic mass and daltons as interchangeable

Coming next

Chapter 3 arranges the elements in the periodic table: groups and periods, metals and non-metals, and how the table's layout follows from the structure of the atom.

Chemistry Fundamentals

Course 1 · Chapter 3 · The Periodic Table: A Map of the Elements

On 6 March 1869 the Russian chemist Dmitri Mendeleev presented a paper to the Russian Chemical Society, and according to Wikipedia's article on him it was titled "The Dependence between the Properties of the Atomic Weights of the Elements". Its central claim was that elements, "if arranged according to their atomic weight, exhibit an apparent periodicity of properties". That sentence is the idea behind the periodic table: the chemical elements are not a random list but fall into a pattern. This chapter shows how the modern table is laid out, how to sort its elements into metals, non-metals and metalloids, what three well-known families have in common, how Mendeleev used his pattern to predict elements nobody had found, and how to read one element's box. It builds on Chapter 2, where the atomic number was defined as the number of protons.

How the Table Is Laid Out

Wikipedia's Periodic table article describes it as "an ordered arrangement of the chemical elements into rows ("periods") and columns ("groups")", with the elements "listed in order of increasing atomic number". So the atomic number from Chapter 2 is the ordering key: hydrogen (1) comes first, and each step along the sequence adds one proton.

Period (row)

A horizontal row. The article lists periods 1 to 7 (with an extended "8+" entry for elements not yet made), and says a new period "is started when a new electron shell has its first electron". What an electron shell is comes in a later chapter.

Group (column)

A vertical column. The article says groups are numbered 1 to 18 under the IUPAC system, with the f-block groups ignored in that numbering.

Block

"The table is also divided into four roughly rectangular areas called blocks." The article names them s, p, d and f.

The same article says: "Today, 118 elements are known, the first 94 of which are known to occur naturally on Earth." The remaining 24, "americium to oganesson (95-118)", occur only when synthesised in laboratories. (The subtraction 118 - 94 = 24 was run in Python and agrees with the page.) This agrees with the figure of 118 elements used in Chapter 1.

> **What I did and did not read**
>
> The Periodic table page is about 250,800 characters long and I read only the first 100,000. The sections on the table's history beyond Mendeleev, on its many alternative forms and on the placement debates fell mostly in the unread part, so this chapter does not cover them.

Why a Group Behaves Alike

The Periodic table article says: "Elements with similar chemical properties generally fall into the same group in the periodic table." That is the practical value of the table: knowing one element tells you a lot about the elements above and below it. For the reason, the article points to electron arrangement: elements with the same number of electrons in a particular subshell share a column. The Alkali metal article makes the same point for group 1, saying the elements all have an outermost electron in an s-orbital and that "this shared electron configuration results in them having very similar characteristic properties".

The explanation in terms of electrons needs ideas (shells, subshells, orbitals) that this course reaches in a later chapter, so for now treat it as a stated fact: same group, similar chemistry, and the cause lies in the electrons.

Metals, Non-metals and Metalloids

The Periodic table article describes two trends: "Metallic character increases going down a group and from right to left across a period", and "Nonmetallic character increases going from the bottom left of the periodic table to the top right". Wikipedia's Nonmetal article agrees on position: nonmetals "nearly all occupy the right-hand side of the periodic table".

Class	What the pages say	Source
Metals	Mostly the left and centre of the table. The Alkali metal article calls its metals "shiny, soft, highly reactive". The Nonmetal page gives no count of metal elements.	Periodic table; Alkali metal; Nonmetal
Non-metals	"A chemical element that mostly lacks distinctive metallic properties"; "usually less dense than metals and are often poor conductors of heat and electricity"; solids are "brittle or crumbly rather than malleable or ductile". "Seventeen elements are widely recognized as nonmetals", and "about half" are gases near standard conditions; bromine is the only liquid.	Nonmetal
Metalloids	"The six commonly recognised metalloids are boron, silicon, germanium, arsenic, antimony and tellurium." They are "shiny, brittle solids with intermediate to relatively good electrical conductivity". "Silicon is the leading commercial semiconductor."	Metalloid

The boundaries are fuzzy

The Metalloid article says no single accepted definition exists and that experts disagree about which elements qualify. The Nonmetal article itself gives two counts, seventeen "widely recognized" and fourteen "almost always recognized", depending on how the borderline cases are treated. So "how many non-metals are there" has no single answer, and a question like that should always name its source.

Pages read: Metalloid and Nonmetal were each about 234,000 and 259,000 characters long, and I read the first 100,000 of each. Hydrogen is a special case: the Alkali metal article says hydrogen is included in group 1 ("together with hydrogen they constitute group 1"), though hydrogen is not itself one of the alkali metals it lists. That hydrogen is a non-metal is my general knowledge (check), and I did not read further into why it is placed there.

Three Families

Alkali metals (group 1)

From the Alkali metal article: lithium, sodium, potassium, rubidium, caesium and francium, with hydrogen sharing the group. They are "shiny, soft, highly reactive metals" that "can all be cut easily with a knife". "All the alkali metals react with water, with the heavier alkali metals reacting more vigorously than the lighter ones." Because of this they are "usually stored in mineral oil or kerosene (paraffin oil)".

Alkali metal	Density (g/cm^3)	Melting point (°C)
Lithium	0.534	180.54
Sodium	0.968	97.72
Potassium	0.89	63.38

Going down the group, the page lists increasing reactivity and decreasing melting point alongside increasing atomic radius. The table shows the melting-point trend. The densities do not fall in a simple order (sodium is higher than potassium), so not every property is a smooth trend. I read the first 100,000 of about 295,000 characters of this page.

Halogens (group 17)

The Halogen article says: "In the modern IUPAC nomenclature, this group is known as group 17." It lists fluorine, chlorine, bromine and iodine, plus radioactive astatine and tennessine (some authors exclude tennessine). The name means "salt former" or "salt maker", and the page says that halogens reacting with metals produce salts such as sodium chloride and potassium iodide. One picture caption states: "Chlorine is a gas, bromine is a liquid, and iodine is a solid." The page says "the reactivity of elements decreases because of the increasing size of the atoms", so reactivity falls going down the group, the opposite of the alkali metals.

Halogen	Melting point (°C)	Boiling point (°C)
Chlorine	-101.5	-34.04
Bromine	-7.3	58.8
Iodine	113.7	184.3

I read about 100,000 of 207,600 characters of this page.

Noble gases (group 18)

The Noble gas article calls them "the members of group 18". Under standard conditions they are "colorless, odorless, tasteless, and nonflammable", and the first six are "monatomic gases with very low chemical reactivity" (monatomic means one atom per particle). The page attributes the low reactivity to a full outer electron shell. Boiling points from its table are helium 4.4 K, neon 27.3 K and argon 87.4 K. Older layouts, the page says, labelled this column group 0 or group 8. Uses it lists include helium in balloons, argon shielding welding and glowing discharge lamps such as neon signs. I read about 100,000 of 215,600 characters.

> **One pattern, two directions**
>
> Reactivity rises down group 1 and falls down group 17, so "down the group" does not always mean "more". Always check a specific group's trend before generalising. Noble gases barely react at all.

Mendeleev's Table and His Predictions

Wikipedia's Periodic table article says: "The first periodic table to become generally accepted was that of the Russian chemist Dmitri Mendeleev in 1869." He framed the periodic law around atomic mass, and because some elements were still unknown, he used the law to predict properties of the missing ones. The Dmitri Mendeleev article says he arranged elements by atomic weight (now called relative atomic mass, as in Chapter 2) and valence, and that his predicted elements were called "ekasilicon, ekaaluminium and ekaboron", corresponding to germanium, gallium and scandium. Gallium and germanium were found in "1875 and 1886 respectively". The Wikipedia article on Mendeleev's predicted elements explains that the prefix "eka" comes from a Sanskrit word meaning "one", and that Mendeleev used eka-, dvi- or tri- depending on how many places down the group the missing element lay below a known one.

Property	Predicted (ekasilicon)	Observed (germanium)	Predicted (ekaaluminium)	Observed (gallium)
Atomic mass	72	72.630	68	69.723
Density (g/cm^3)	5.5	5.323	6.0	5.91

Property	Predicted (ekasilicon)	Observed (germanium)	Predicted (ekaaluminium)	Observed (gallium)
Melting point	"High"	938 °C	"Low"	29.76 °C

All values are from the Mendeleev's predicted elements page. The differences (for example 72.630 - 72 = 0.63 and 6.0 - 5.91 = 0.09) are my own arithmetic in Python. The page also says gallium was discovered "soon after Mendeleev predicted its existence, in 1875 by Paul Emile Lecoq de Boisbaudran". The Mendeleev article gives no discovery year for scandium and no predicted property values, so this chapter gives none. It adds that Mendeleev was wrong to assume atomic weight must always increase in order: for tellurium he said the weight "must lie between 123 and 126", while the observed value is 127.6.

The Periodic table article also notes that "Group 18, the noble gases, had not been discovered at the time of Mendeleev's original table", so the table had a column that did not exist yet. I read all of the Mendeleev's predicted elements page and 100,000 of about 151,300 characters of the Mendeleev biography.

Reading One Element's Box

The Periodic table article says each tile shows the symbol and a standard atomic weight, with the atomic number printed in front of the symbol. Its example reads "Hydrogen 1 H 1.0080". Reading it:

- **1** is the atomic number: the number of protons (Chapter 2).
- **H** is the symbol.
- **1.0080** is the standard atomic weight, the weighted average over isotopes described in Chapter 2 as relative atomic mass.

The article also says that a bracketed value such as "[209]" is the mass number of the most stable isotope: where an element has no stable isotopes, there is no sensible average, so the table gives one isotope's mass number instead. The relation N = A - Z from Chapter 2 then gives that isotope's neutron count (a worked case is in Exercise 1).

Where to Read More

- **Encyclopaedias.** The Wikipedia articles Periodic table, Alkali metal, Halogen, Noble gas, Metalloid, Nonmetal, Dmitri Mendeleev and Mendeleev's predicted elements are the sources used here. Apart from the predicted-elements page, each was read only up to its first 100,000 characters, as noted above.
- **A school textbook.** A printed periodic table shows every box for all 118 elements.
- **Practical.** Take a periodic table and colour in the metalloids, then check how they follow a rough diagonal line.

Exercises

EXERCISE 1

Read an element's box. From the Hydrogen example in this chapter, give its atomic number, symbol and atomic weight, and the number of protons and electrons in a neutral hydrogen atom. Then suppose a box shows a bracketed [209] for an element whose atomic number is 84 (the 84 is given for this exercise; check it against a periodic table). What does the bracket mean, and how many neutrons does that isotope have?

→ Solution

EXERCISE 2

Classify using the pages' statements. (a) Name the group number and family of chlorine, of potassium and of argon. (b) Which of boron, silicon, sodium, germanium and bromine does the Metalloid article list as commonly recognised metalloids? (c) Of lithium, sodium and potassium, which is expected to be the most reactive with water, and which the most metallic in character? Give the statement that supports each answer.

→ Solution

EXERCISE 3

Estimate, as Mendeleev's method suggests, from neighbours. Using only the melting points of chlorine (-101.5 °C) and iodine (113.7 °C) from this chapter, estimate the

melting point of bromine by taking their average. Compare with the page's value of -7.3 °C, say how far off the estimate is, and explain why this method is only a rough guide.

→ Solution

Quick Reference

- **Layout:** periods are rows (1 to 7), groups are columns (1 to 18); elements run in order of atomic number
- **Elements:** 118 known; the first 94 occur naturally; 95 to 118 are made in laboratories
- **Trends:** metallic character rises down a group and right to left across a period
- **Metalloids:** boron, silicon, germanium, arsenic, antimony, tellurium (the six commonly recognised)
- **Families:** group 1 alkali metals (reactivity rises down); group 17 halogens (reactivity falls down); group 18 noble gases (very low reactivity)
- **Mendeleev:** 1869 table by atomic weight; predicted ekasilicon, ekaaluminium and ekaboron (germanium 1886, gallium 1875, scandium)
- **Box:** atomic number, symbol, atomic weight; a bracketed number is the mass number of the most stable isotope
- **Check:** the atomic number 84 used in Exercise 1; the sorting of unnamed borderline cases

Coming next

Chapter 4 covers ions and ionic bonding: electron transfer, ionic compounds, formulae and lattices, and begins to use the electron arrangements behind the groups.

Chemistry Fundamentals

Course 1 · Chapter 4 · Ions and Ionic Bonding: Atoms That Carry a Charge

Table salt is sodium chloride, and it is made of two elements that on their own are very different: Chapter 3 described sodium as a soft, highly reactive alkali metal and chlorine as a halogen gas. Wikipedia's Ionic bonding article defines the link between them: "Ionic bonding is a type of chemical bonding that involves the electrostatic attraction between oppositely charged ions." This chapter explains what an ion is, how atoms become ions by gaining or losing electrons, why the metals and non-metals of Chapter 3 tend to go different ways, how to write a formula by balancing charges, and what ionic compounds are like. It builds on Chapter 2, where an atom's protons and electrons were counted, and it is the first chapter to use electrons to explain behaviour, a promise made at the end of Chapter 3.

What an Ion Is

According to Wikipedia's Ion article, an ion is "an atom or molecule with a net electrical charge". The charge appears because the counts of protons and electrons no longer match: "The net charge of an ion is not zero because its total number of electrons is unequal to its total number of protons." In Chapter 2 a neutral atom had equal numbers of the two, so an ion is what you get when that balance is broken by gaining or losing electrons. The protons stay put; only electrons move.

Cation

The article defines a cation as "a positively charged ion with fewer electrons than protons". Losing electrons makes an atom positive.

Anion

An anion is "a negatively charged ion with more electrons than protons". Gaining electrons makes an atom negative.

Writing the charge

The article says singly charged ions are marked with just a + or a −, such as Na+ and F−. A charge of two or three is written

with the number, as in 3+ and 2− later in this chapter.

A way to remember which is which

This memory aid is my own, not from the pages (check): the charge sign tells you what happened to the electrons. A plus sign means electrons were lost (fewer electrons than protons), and a minus sign means electrons were gained (more electrons than protons).

Valence Electrons: Which Electrons Move

Not all of an atom's electrons take part. Wikipedia's Valence electron article defines them: "valence electrons are electrons in the outermost shell of an atom", and says they can take part in bonding "if the outermost shell is not closed". It also gives a rule for reading the count from the periodic table: "In groups 1-12, the group number matches the number of valence electrons", and for groups 13 to 18 "the units digit of the group number matches the number of valence electrons", with helium as the exception. So by that rule sodium (group 1) has one valence electron and chlorine (group 17) has seven. That is the electron explanation of group behaviour that Chapter 3 put off: atoms in one column have the same number of outer electrons, so they behave alike.

Why do atoms gain or lose them? The Octet rule article describes "the strong tendency for main-group elements to bond in such a way that each atom has eight electrons" in their outer shell. The Valence electron article connects this to the noble gases of Chapter 3: an atom with a closed outer shell tends to be chemically inert, which fits the low reactivity noted for group 18. Sodium "easily" loses its single valence electron "to form a positive ion", while chlorine "requires only one additional valence electron to form a closed shell". Shells beyond this point (what a shell holds, subshells, orbitals) are left for a later look; here the rule of thumb is enough.

What I did and did not read

The Ion, Ionic bonding, Valence electron and Octet rule pages were fetched in full or in the part the tool returned, and the tool did not report their lengths, so I cannot say what fraction of each I read. The Ionic compound page is 139,910 characters and I read the first

100,000. The Sodium chloride page is 122,180 characters and I read the first 100,000. Anything not tied to one of these pages is marked "(check)".

Sodium Chloride: The Worked Example

The Ionic bonding article walks through this case: "the sodium atoms each lose an electron, forming cations (Na+)", and chlorine atoms gain those electrons and become anions, Cl−. The Octet rule article says it directly: "Sodium is observed to transfer one electron to chlorine during the formation of sodium chloride", and "Adding an electron to chlorine produces the chloride ion (Cl−)". The Valence electron article agrees that in solid sodium chloride the sodium valence electron "is transferred to chlorine to form an ionic bond". The two ions attract in a 1:1 ratio, giving NaCl.

Step	Sodium	Chlorine
Group (Chapter 3)	1, an alkali metal	17, a halogen
Valence electrons (group rule)	1	7
What happens	loses 1 electron	gains 1 electron
Result	Na+ (cation)	Cl− (anion)

Counting protons and electrons (Chapter 2) shows the charge. Sodium's atomic number is 11 and chlorine's is 17 (check, these are from general knowledge, not a page I read in this chapter). A neutral sodium atom has 11 protons and 11 electrons; after losing one, 11 protons and 10 electrons remain, and 11 - 10 = +1. A chlorine atom has 17 and 17, and after gaining one it has 17 protons and 18 electrons, so the net charge is 17 - 18 = -1. (The subtractions 11 - 1 = 10 and 17 + 1 = 18 were run in Python.) These counts match the Na+ and Cl− charges given by the pages. The pages do not themselves give the atomic numbers, so they are shown as unverified.

Why Metals Form Cations and Non-metals Form Anions

The Ion article says it plainly: "metals will lose electrons to form positively charged ions", while non-metals gain electrons to form negatively charged ones. The Ionic bonding article adds that "in the simplest case, the cation is a metal atom and the anion is a nonmetal atom". This links to

Chapter 3: metals lie to the left and centre of the table, non-metals toward the right, and the halogens just before the noble gases. Chlorine, one electron short of a full outer shell, takes one. Sodium, one electron beyond a closed shell, gives one away. The pages do not explain in energy terms why sodium finds this easier than chlorine does; that is a deeper question left alone here.

> **Transfer, not just a loss**
>
> The electron lost by sodium does not vanish. In sodium chloride it ends up on chlorine, so the total number of electrons is unchanged and the compound as a whole has no net charge. The Ionic compound article states that such a compound has "no net electric charge".

Ionic Bonding as Attraction

The bond itself is the pull between opposite charges. The Ion article says "cations and anions attract each other and readily form ionic compounds", and the Ionic bonding article calls the bond "the electrostatic attraction between oppositely charged ions". The Ionic compound article says the ions are held together "by electrostatic forces called ionic bonds". Electrons are not shared here; they are handed over. Chapter 5 turns to the other main case, where atoms share electrons.

Writing a Formula from Charges

The Ionic bonding article gives the rule: "to maintain charge neutrality, strict ratios between anions and cations are observed." So a formula is found by choosing the smallest numbers of each ion that make the total charge zero. For sodium chloride, +1 and -1 cancel at 1:1, giving NaCl, as the Sodium chloride article's formula and "1:1 ratio" show.

A case with unequal charges comes from the Ionic compound article, which says that in $Fe_2(SO_4)_3$ each iron ion has a 3+ charge, balancing sulfate ions at 2− each. Check the balance: two iron ions give 2 x 3 = +6, three sulfate ions give 3 x 2 = -6, and the total is 6 - 6 = 0 (all run in Python). Notice how the subscripts come from the charges: the number of each ion needed is found by making the totals equal, and the sulfate ion is written in brackets because it is a group of atoms carrying the charge together. The article describes such ions as polyatomic.

Compound	Ions	Charge check
Sodium chloride, NaCl	Na+ and Cl−	(+1) + (-1) = 0
Iron sulfate, $Fe_2(SO_4)_3$	Fe 3+ and SO_4 2−	2 x (+3) + 3 x (-2) = 0

A method for any pair

The method is my own summary (check): write the two charges, ignore the signs, then swap the numbers to become the subscripts and divide both by any common factor. It matches the cases above (3 and 2 swap to give 2 iron and 3 sulfate). Always confirm by adding the charges and getting zero, as the pages require.

Properties of Ionic Compounds

Several properties follow from the picture of charged ions held by strong attraction. The pages state the following.

Property	What the page says	Source
Melting point	"Ionic compounds generally have a high melting point", rising with the ion charges. Sodium chloride melts at 801 °C (the infobox lists 800.7 °C).	Ionic bonding; Sodium chloride
Boiling point	Salts of small ions "typically have high melting and boiling points". Sodium chloride is given as 1465 °C in the text but 1,413 °C in the infobox, so the page disagrees with itself by 52 °C (Python).	Ionic compound; Sodium chloride
Electrical conduction	"Ionic compounds conduct electricity when molten or in solution, typically not when solid." The Ionic compound page says solids are "almost always electrically insulating" and become "highly conductive" when melted or dissolved, because the ions become mobile.	Ionic bonding; Ionic compound
Brittleness	"Most salts are very brittle."	Ionic compound

Property	What the page says	Source
Dissolving	"When simple salts dissolve, they dissociate into individual ions." Solubility is highest in polar solvents such as water.	Ionic compound

The conduction rule is a good way to remember what is going on. Electric current through a liquid is carried by charged particles that are free to move. In the solid the ions are locked in place, so there is no current; melting or dissolving lets them move. Spelling out that reason is my own wording (check), though the pages give the same cause for the Ionic compound page's conductivity statement. The pages do not say why brittleness occurs, so this chapter does not explain it.

A Lattice, Not a Molecule

A water molecule is a small, countable unit. Ionic compounds are different. The Ionic bonding article says ionic solids form extended lattices, so "it is usually not possible to distinguish discrete molecular units". The Ionic compound article calls the ions "part of a continuous three-dimensional network". The Sodium chloride article gives the real arrangement: it "can be represented as a face-centered cubic (fcc) lattice with a two-atom basis", and "each ion is surrounded by six ions of the opposite charge". So a grain of salt is one huge array in which every Na+ sits among six Cl− and every Cl− among six Na+, not a heap of separate NaCl molecules.

That is why NaCl is called a formula, not a molecular formula: it gives the simplest ratio of ions. The Ionic compound article uses the phrase "formula unit" for one such group, as in "the two iron ions in each formula unit" of $Fe_2(SO_4)_3$. I did not read how that article defines the term, so treat "formula unit" as the smallest repeating ratio (check).

> **Where this leaves us**
>
> Ionic bonding explains one kind of compound well: metal plus non-metal, electrons handed over, a lattice of ions, a high melting point, conduction only when the ions can move. Many substances do not fit this pattern, and Chapter 5 covers those.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles Ion, Ionic bonding, Ionic compound, Valence electron, Octet rule and Sodium chloride are the sources used here. See the note above for how much of each was read.
- **A school textbook.** Look for a table of common ions and their charges, which this chapter does not reproduce. Ion also has a table of common anions that I did not quote.
- **Practical.** Dissolve a little salt in a glass of water and look up how a conductivity test would differ from pure water, then compare with the page's statement on solutions.

Exercises

EXERCISE 1

Predict an ion from position. Using the group rule from the Valence electron article, give the number of valence electrons of potassium (group 1) and fluorine (group 17). Say whether each is more likely to lose or gain electrons, how many, and write the ion with its charge. State which is the cation and which the anion, and how the number of electrons compares with the number of protons in each ion (use a general atomic number of your own and mark it as such).

→ Solution

EXERCISE 2

Balance a formula from charges. (a) Using the Fe 3+ and sulfate 2− charges given in this chapter, show that $Fe_2(SO_4)_3$ has zero total charge. (b) For this exercise only, suppose an ion X has charge 2+ and an ion Y has charge 1−. Find the formula of the compound by balancing the charges, and show the check. (c) Suppose instead X has charge 3+ and Y has charge 2−: find the formula and check it.

→ Solution

EXERCISE 3

Decide from properties. Substance P is a hard, brittle solid that melts at a high temperature. It does not conduct electricity as a solid, but it does when melted and

when dissolved in water. Substance Q is a soft solid that conducts electricity as a solid. Which, if either, matches the pages' description of an ionic compound, and which statements support the decision? Then say why a lattice picture, not a molecule picture, suits your answer, and what the pages do not tell you about substance P.

→ Solution

Quick Reference

- **Ion:** an atom or molecule with a net electrical charge; protons and electrons no longer match
- **Cation / anion:** positive (fewer electrons than protons) / negative (more electrons than protons)
- **Valence electrons:** outer-shell electrons; for groups 1 to 12 the group number, for groups 13 to 18 the units digit (helium excepted)
- **Octet rule:** main-group atoms tend to bond so each has eight outer electrons
- **Sodium chloride:** Na loses one electron to become Na+, Cl gains one to become Cl−, 1:1 ratio, NaCl
- **Who forms what:** metals lose electrons to form cations; non-metals gain electrons to form anions
- **Bond:** electrostatic attraction between oppositely charged ions
- **Formula:** choose ion numbers so the total charge is zero, as in $Fe_2(SO_4)_3$: 2 x 3 = 6 and 3 x 2 = 6
- **Properties:** high melting points, brittle, conduct when molten or dissolved but typically not as solids
- **Structure:** a three-dimensional lattice; in sodium chloride each ion has six neighbours of opposite charge
- **Check:** the atomic numbers of sodium and chlorine; the reading of "formula unit"; my own memory aids

Coming next

Chapter 5 turns to covalent bonds and molecules: what happens when atoms share electrons instead of handing them over, and how to draw and read the result.

Chemistry Fundamentals

Course 1 · Chapter 5 · Covalent Bonds and Molecules: Atoms That Share

Chapter 4 described atoms that hand electrons over and then cling together as oppositely charged ions. Plenty of familiar substances are not like that. The oxygen you breathe is an element, as Chapter 1 listed it, yet it is not a crowd of separate atoms: it is made of atoms joined in pairs. Wikipedia's Covalent bond article defines the kind of link that does the joining: "a chemical bond that involves the sharing of electrons to form electron pairs between atoms". This chapter covers sharing instead of transferring, how to count and draw shared pairs, why some bonds pull harder to one side than the other, and why the shape of a molecule matters. It uses the valence electrons introduced in Chapter 2 and the group rule from Chapter 4.

Sharing Instead of Handing Over

The Covalent bond article calls the shared electron pairs "shared pairs" or "bonding pairs". It explains why atoms do this with Lewis's idea: "Lewis proposed that an atom forms enough covalent bonds to form a full (or closed) outer electron shell." That is the octet rule of Chapter 4 again, reached by a different route. Its example is hydrogen: "the hydrogen atoms share the two electrons via covalent bonding", so each hydrogen atom in an H_2 unit has two electrons around it, a full first shell (the full first shell of two is my addition, not a quote from the page (check)).

Ionic (Chapter 4)

Electrons move from one atom to another. The Ionic bonding article calls the bond an electrostatic attraction between oppositely charged ions.

Covalent (this chapter)

Electrons are shared between atoms as bonding pairs. The Covalent bond article contrasts it with ionic bonding, where atoms "are separate and merely associated by electrostatic attraction".

Which elements?

The pages I read do not give a simple rule. A common textbook shortcut is that two non-metals usually share (check), and Chapter 4

said that in the simplest ionic case the cation is a metal and the anion a non-metal.

Single, Double and Triple Bonds

Atoms can share more than one pair. The Covalent bond article describes the three cases: "A single bond is usually a σ bond", "a double bond between two given atoms consists of one σ and one π bond", and "a triple bond is one σ and two π bonds". The Greek-letter names describe how the electron clouds overlap, which this course does not go into. What matters here is the count: a single bond is one shared pair, a double bond two pairs, a triple bond three. Counting pairs gives the **bond order**. The page has no general definition of the term, but for oxygen it writes "The bond order is 1+0.5+0.5=2", and the Diatomic molecule article says that diatomic oxygen has a double bond.

Bond	Shared pairs	Shared electrons	Example
Single	1	2	H_2 (the page: hydrogen atoms share two electrons)
Double	2	4	O_2 (the Diatomic molecule article: "diatomic oxygen has a double bond")
Triple	3	6	N_2 (a triple bond is my own recall, not read on a page (check))

The shared-electron column is my own arithmetic, two electrons per pair, run in Python: 1 x 2 = 2, 2 x 2 = 4 and 3 x 2 = 6.

Molecules and Molecular Formulae

Wikipedia's Molecule article defines a molecule as "a group of two or more atoms that are held together by attractive forces known as chemical bonds". A molecular formula "reflects the exact number of atoms that compose the molecule". The article's examples are water, "two hydrogen atoms and one oxygen atom; H2O", and the oxygen molecule, "two atoms in the oxygen molecule (O2)". The subscript counts atoms of the element in front of it, and no subscript means one. This is the difference from Chapter 4: the article says "No typical molecule can be defined

for salts", and that atoms joined by ionic bonds "are typically not considered single molecules". Sodium chloride was a lattice with a formula ratio; water is a countable unit of three atoms.

> **Not every molecule has two atoms**
>
> The same article notes that in the kinetic theory of gases the word is used more loosely, "since the noble gases are individual atoms". In the strict definition above a molecule has two or more atoms; a lone atom of a noble gas from Chapter 3 is the loose usage.

Diatomic Elements: The Oxygen Promise

Chapter 1 gave oxygen gas as an example of an element, which it defined as a substance made of one kind of atom. Here is how that fits with atoms and molecules. The Diatomic molecule article says such molecules "are molecules composed of only two atoms, of the same or different chemical elements". It lists the elements that are diatomic at 1 bar and 25 °C: "the gases hydrogen (H2), nitrogen (N2), oxygen (O2), fluorine (F2), and chlorine (Cl2), and the liquid bromine (Br2)". It adds that iodine forms a diatomic gas at slightly higher temperatures. So the oxygen in air is O_2: two oxygen atoms joined by a double bond, still one element because both atoms are the same kind. Chlorine is on the list too: the chlorine atom that took an electron from sodium in Chapter 4 is not how chlorine exists on its own, which is Cl_2 (the link to Chapter 4 is my own (check)).

> **Element does not mean single atoms**
>
> Chapter 1's definition still holds: oxygen gas is an element because it contains only oxygen atoms. What this chapter adds is that the atoms come in joined pairs. Counting the pair as one particle is what a molecular formula does: O_2, not O.

Drawing a Lewis Structure

The Lewis structure article describes these as "diagrams that show the bonding between atoms of a molecule", where "excess electrons that form lone pairs are represented as pairs of dots" and "lines are drawn between atoms that are bonded to one another". The page's own full worked example is the nitrite ion, and it shows a water molecule only as a picture caption, so the

steps below are my own application of the page's rules, using the valence-electron group rule from Chapter 4 (check).

Step	For water, H_2O
1. Count valence electrons	Oxygen is group 16, so 6; each hydrogen is group 1, so 1. Total 6 + 1 + 1 = 8 (Python).
2. Join atoms with single bonds	Oxygen in the middle, one line to each hydrogen. Two bonds use 2 x 2 = 4 electrons.
3. Place the rest as lone pairs	8 - 4 = 4 electrons left, which is 2 lone pairs, placed on oxygen.
4. Check the shells	Oxygen has 2 bonds (4 electrons) plus 2 lone pairs (4 electrons) = 8. Each hydrogen shares one pair = 2.

```
    ..
H — O — H
    ..

Each line   = one shared pair (2 electrons)
Each ..     = one lone pair (2 electrons)
Total       = 2 bonds x 2 + 2 lone pairs x 2 = 8 electrons
```

Everything in the drawing is accounted for: the 8 valence electrons counted in step 1 equal the 8 electrons drawn. Counting back like this is the best check that a Lewis structure is right. Note that the drawing shows which atoms are joined, not the angle between the bonds; the shape section below comes to that.

Electronegativity and Polar Bonds

In a shared pair, the two atoms do not necessarily pull equally. Wikipedia's Electronegativity article defines electronegativity as "the tendency for an atom of a given chemical element to attract shared electrons". The Covalent bond article says covalent character "is greatest between atoms that have similar electronegativities", and "an unequal relationship creates a polar covalent bond such as with H−Cl". Equal sharing, as in H–H, gives a nonpolar bond.

The Chemical polarity article gives a rough guide: polar bonds "generally occur when the difference in electronegativity between the two atoms is roughly between 0.5 and 2.0", with

smaller differences nonpolar and larger ones ionic. The Electronegativity article itself gives no numeric cut-offs and says bonding falls along a continuous scale from covalent to ionic, so treat the numbers as a guide, not a law. Below are Pauling-scale values from the Electronegativity article, and the differences I calculated in Python.

Bond	Values (Electronegativity article)	Difference (Python)	By the guide
H–H	2.20, 2.20	0.00	nonpolar
C–H	2.55, 2.20	0.35	nonpolar
O–H	3.44, 2.20	1.24	polar
H–Cl	2.20, 3.16	0.96	polar
Na–Cl	0.93, 3.16	2.23	ionic

The last row ties back to Chapter 4: the sodium chloride example sits on the ionic side of the guide, and the article's idea of a continuous scale means it is not a sharp line. In a polar bond the more electronegative atom is at the negative end of the dipole, so in H–Cl the chlorine end is the negative one (the Electronegativity article).

> **A bond can be polar in a molecule that is not**
>
> The Chemical polarity article makes the point with carbon dioxide: its C=O bonds are polar, but "the geometry of CO2 is linear so that the two bond dipole moments cancel", so the molecule is nonpolar. It also says methane's "bonds are arranged symmetrically (in a tetrahedral arrangement) so there is no overall dipole". Polarity of a bond and polarity of a molecule are two separate questions, and shape decides the second.

Shape: Why Water Is Bent

Water is the standard example. The Properties of water article gives its shape as "bent", not linear, and the Chemical polarity article says the bent geometry means the O–H bond dipoles do not cancel, so the molecule has a net dipole and is polar. The Properties of water article states the angle twice: "The hydrogen–oxygen–hydrogen angle is 104.45°" in one place, and a gas-phase bend angle of 104.48° in another, which differ by 0.03° (Python). So say "about 104.5°"

and note the page is not consistent to the second decimal place. The methane article calls the molecule "a tetrahedral molecule with four equivalent C–H bonds"; it did not give the bond angle in the part I read, so no angle is given here.

> **Why the Lewis drawing can mislead**
>
> The text drawing of water above shows a straight line, H–O–H. The real molecule is bent at about 104.5°. The Lewis structure tells you which atoms are bonded and where the lone pairs are, not the three-dimensional shape. Predicting shape from the electron pairs is a separate method that this chapter does not cover (check; the VSEPR theory page was not read).

Ionic and Covalent Compounds Compared

Chapter 4 gave the properties of ionic compounds. The Covalent bond article describes covalent substances in two groups and does not compare them directly with ionic ones, so the table below sets the two side by side using only wording from the pages I read; a blank cell means the pages did not say.

Feature	Ionic (Chapter 4 pages)	Covalent (Covalent bond article)
Bond	Electrostatic attraction between oppositely charged ions	Sharing of electron pairs between atoms
Basic unit	A lattice, "usually not possible to distinguish discrete molecular units"	Molecules in simple molecular substances, or giant network structures
Melting and boiling	Generally high melting point	Molecular substances: low-boiling liquids and low-melting solids; network structures: "high melting and boiling points"
Forces between particles	Strong electrostatic forces	Molecular substances: "negligible forces of attraction between molecules" (the page's wording)

Feature	Ionic (Chapter 4 pages)	Covalent (Covalent bond article)
Electricity	Conduct when molten or dissolved, typically not as solids	Network structures "tend to have high electrical resistivity"; nothing stated for molecular substances
Brittleness	"Most salts are very brittle"	Network structures "frequently brittle"
State at room temperature		"Such covalent substances are usually gases" (simple molecular substances)

The table shows that "covalent" covers two quite different kinds of material, and that one set of properties does not fit them all. The pages give no melting points for network solids, so the comparison stays qualitative. For two molecular examples, the Properties of water article gives 0.00 °C for melting and 99.98 °C for boiling; the Methane article gives −182.456 °C and −161.49 °C. Compare these with sodium chloride's 801 °C melting point from Chapter 4. Three substances are a small sample, not proof of a pattern.

What I did and did not read

The Covalent bond, Diatomic molecule, Lewis structure and Chemical polarity pages appeared complete in what the tool returned. The Molecule page is 118,700 characters and I read the first 100,000. The Electronegativity page is 114,126 characters and I read the first 100,000. The Properties of water page is 241,084 characters and I read the first 100,000. The Methane page is 208,623 characters and I read the first 100,000. Statements marked "(check)" come from my own knowledge or reasoning. The electronegativity values and other quotations came through the fetch tool's summary of each page, so check them against a data table before relying on them.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles Covalent bond, Molecule, Diatomic molecule, Lewis structure, Electronegativity, Chemical polarity, Properties of water and Methane are the sources used here. See the note above for how much of each was read.
- **A school textbook.** Look for worked Lewis structures of ammonia, carbon dioxide and nitrogen, and for the VSEPR method of predicting shape, which this chapter leaves out.

- **Practical.** A molecular model kit (balls and sticks) makes the bent shape of water and the tetrahedral shape of methane easy to see.

Exercises

EXERCISE 1

Count the shared pairs. For each of H_2, O_2 and Cl_2, say how many electron pairs the two atoms share if the bond is single, double and single respectively, give the bond order, and work out the number of shared electrons. Then say which of the three the pages explicitly describe as having that bond order, and which you are assuming.

→ Solution

EXERCISE 2

Draw a Lewis structure. Using the group rule for valence electrons (hydrogen group 1, carbon group 14, chlorine group 17), work out the total for hydrogen chloride, HCl, and draw its Lewis structure with a line for the bond and dots for lone pairs. Check that the electron count matches and that each atom has a full outer shell. Then do the same for methane, CH_4.

→ Solution

EXERCISE 3

Polar or not? Using the Pauling values in this chapter (H 2.20, C 2.55, N 3.04, O 3.44, F 3.98, Na 0.93, Cl 3.16), calculate the electronegativity difference for N–H, H–F and C–O, and classify each as nonpolar, polar or ionic by the guide in the Chemical polarity article. Then explain why a molecule with polar bonds can still be nonpolar, using the carbon dioxide example from the chapter.

→ Solution

Quick Reference

- **Covalent bond:** sharing of electrons to form electron pairs between atoms
- **Single / double / triple:** one / two / three shared pairs, so bond order 1 / 2 / 3
- **Molecule:** two or more atoms held together by chemical bonds; the formula gives the exact number of each atom
- **Diatomic elements:** H_2, N_2, O_2, F_2, Cl_2 (gases), Br_2 (liquid), at 1 bar and 25 °C
- **Lewis structure:** lines for bonded pairs, dots for lone pairs; count the valence electrons first and check the total
- **Electronegativity:** the tendency of an atom to attract shared electrons; a bigger difference gives a more polar bond
- **Guide:** difference below 0.5 nonpolar, roughly 0.5 to 2.0 polar, above 2.0 ionic (Chemical polarity article; a rough guide)
- **Shape:** water is bent (about 104.5°) and polar; carbon dioxide is linear and nonpolar; methane is tetrahedral
- **Check:** N_2 as a triple bond, the two-non-metals shortcut, the Lewis steps for water, and VSEPR

Coming next

Chapter 6 moves from single molecules to what happens when substances react: how chemical changes are written as equations, and why the numbers of atoms must balance.

Chemistry Fundamentals

Course 1 · Chapter 6 · Chemical Reactions and Equations: Writing Down a Change

Chapter 1 separated physical change from chemical change: in a chemical change new substances form. Chapters 4 and 5 then gave you the formulae of ionic compounds and of molecules. Now the two come together. A chemical equation is the shorthand chemists use to record a chemical change, using those formulae. Wikipedia's Chemical reaction article defines a reaction as "a process that leads to the chemical transformation of one set of chemical substances to another". This chapter covers how to read and write equations, why they must balance, how to balance one step by step, and the main types of reaction. It also covers what an equation leaves out, which is a lot.

Reactants and Products

The Chemical reaction article calls the starting substances "reactants or reagents", and says the reaction yields "one or more products, which usually have properties different from the reactants". That last phrase is Chapter 1's test for chemical change in other words: the product is a new substance with new properties. The Chemical equation article says "The reactant entities are given on the left-hand side, and the product entities are on the right-hand side", and the Chemical reaction article says the arrow "points in the direction in which the reaction proceeds".

Word equation

The substances written in words: hydrogen + oxygen → water. Easy to read, but it says nothing about how many of each (the format is my own teaching device (check)).

Symbol equation

The same change in formulae with coefficients. The Chemical equation article says a coefficient "specifies how many entities (e.g. molecules) of that substance are involved".

State symbols

Small labels after each formula. The Chemical equation article lists "(s) for a solid, (l) for a liquid, (g) for a gas, and (aq) for an aqueous solution".

"Aqueous" means dissolved in water (my own gloss (check)). Take the notation from Chapter 5: O_2 is a molecule of two oxygen atoms, so the subscript sits inside a formula and cannot be changed. The big number in front, the coefficient, is the part you are allowed to adjust.

Why an Equation Must Balance

The Chemical equation article states the requirement directly: "each side of the chemical equation must represent the same number of atoms of any particular element". The reason is the law of conservation of mass. Wikipedia's Conservation of mass article says "the mass of the system must remain constant over time", for systems closed to all matter transfers, and that mass conservation in chemical reactions was "finally confirmed by Antoine Lavoisier in the late 18th century". The page's history section says he "expressed his conclusion in 1773 and popularized the principle of conservation of mass", and that his experiments disproved the phlogiston theory, which held that mass could be gained or lost in combustion. The Chemical reaction article adds a connected credit: "Antoine Lavoisier developed the theory of combustion as a chemical reaction with oxygen."

The picture behind it, using Chapter 2, is that a reaction rearranges atoms and does not create or destroy them. In the lines of the Chemical equation article, the total electric charge also has to match on both sides, which matters once ions are involved (Chapter 4).

> **The caveats on the page**
>
> The Conservation of mass article says the law "holds only in the classical limit", and that mass is not conserved in open systems, though energy transfers are usually too small to measure unless radioactivity or nuclear reactions are involved. It also says Lomonosov's priority is "sometimes challenged", and that the article itself needs more citations. For everyday chemistry, a closed container, balancing is correct. A burning candle seems to lose mass only because the gases escape (my own explanation (check)).

Balancing Step by Step: Hydrogen and Oxygen

The Chemical equation article says simple equations can be balanced "by inspection, that is, by trial and error". It had no hydrogen-and-oxygen example, so the working below is my own, with every count produced by a Python script. Start from the formulae, which must be right first:

hydrogen gas is H_2 and oxygen gas is O_2 (the Diatomic molecule article from Chapter 5), and water is H_2O.

Step	Equation	H (left / right)	O (left / right)
1. Write the unbalanced equation	$H_2 + O_2 \rightarrow H_2O$	2 / 2	2 / 1
2. Fix oxygen: 2 water	$H_2 + O_2 \rightarrow 2H_2O$	2 / 4	2 / 2
3. Fix hydrogen: 2 hydrogen	$2H_2 + O_2 \rightarrow 2H_2O$	4 / 4	2 / 2

Step 1 fails on oxygen: two atoms in, one out. You cannot repair this by writing H_2O_2 or O, because that changes the substance (Chapter 5's molecular formula) and so changes the chemistry; only coefficients may change. Doubling the water fixes oxygen but upsets hydrogen, so a second adjustment follows. This back-and-forth is normal. The Python check of the finished equation gave H = 4 and O = 2 on each side.

```
2H2(g) + O2(g)  →  2H2O(l)

left :  H = 2 x 2 = 4     O = 1 x 2 = 2
right:  H = 2 x 2 = 4     O = 2 x 1 = 2     balanced
```

State symbols were added here from the notation above; that hydrogen and oxygen are gases and water a liquid at ordinary conditions is my own general knowledge (check).

A second example: burning methane

The Chemical equation article and the Combustion article both give methane burning as $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$, and the Combustion article says "When a hydrocarbon burns in oxygen, the reaction will primarily yield carbon dioxide and water." Rebuilding it from the unbalanced start, my Python count gave:

Element	$CH_4 + O_2 \rightarrow CO_2 + H_2O$	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
C	1 / 1	1 / 1
H	4 / 2	4 / 4
O	2 / 3	4 / 4

A balancing routine that usually works

Balance the elements that appear in the fewest substances first, and leave an element that appears alone (such as O_2) until last, because changing its coefficient disturbs nothing else. Then recount every element. This routine is my own advice, not from a page (check).

The Main Types of Reaction

The Chemical reaction article names several types. The quotations are from that page; the example equations are noted where they come from.

Type	What the page says	Example
Synthesis	"two or more simple substances combine to form a more complex substance"	$2H_2 + O_2 \rightarrow 2H_2O$ (my classification (check))
Decomposition	"a more complex substance breaks down into its more simple parts"	$2H_2O \rightarrow 2H_2 + O_2$ (balance checked in Python; the type is my classification (check))
Single displacement	"a single uncombined element replaces another in a compound"	$2HCl + 2Na \rightarrow 2NaCl + H_2$ (the equation is from the Chemical equation article; classing it here is mine (check))
Double displacement	"the anions and cations of two compounds switch places and form two entirely different compounds"	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$ (my example (check))
Combustion	"an element or compound reacts with an oxidant, usually oxygen, often producing energy in the form of heat or light"	$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$ (Combustion article)

All five examples balanced when I ran the atom counts in Python. A reaction can fit more than one label: burning hydrogen is both a synthesis and a combustion if you use the page's definition of combustion (my reading (check)). The labels are a way to recognise patterns and

not exclusive boxes. Note too the Combustion article's definition: "a high-temperature exothermic redox chemical reaction between a fuel (the reductant) and an oxidant"; the terms exothermic and redox are not covered until later.

What a Balanced Equation Does Not Tell You

A balanced equation is a record of the atom count and of the proportions in which substances react. It leaves out a great deal.

- **Whether the reaction happens.** The Chemical reaction article says "Reactions can proceed by themselves if they are exergonic, that is if they release free energy." Writing an equation does not make a reaction occur; the Chemical equation article says that for ionic equations, "If every ion is a 'spectator ion' then there was no reaction."
- **How fast.** The Chemical reaction article says "The speed at which reactions take place is studied by reaction kinetics." The equation has no time in it. Chapter 9 returns to speed.
- **Conditions and catalysts.** The Chemical equation article notes these are shown only if added above or below the arrow.
- **Direction.** The Chemical reaction article says most reactions are reversible, "they can and do run in both directions", and that a double arrow marks equilibrium reactions.
- **Masses.** A coefficient counts molecules, not grams. Turning counts into amounts you can weigh is the job of Chapter 7.

Common mistakes

Changing a subscript to make the numbers fit (this turns water into something else); forgetting that a coefficient multiplies every atom in the formula it stands before, so $2H_2O$ has 4 hydrogen and 2 oxygen; and balancing the atoms but forgetting to recount after the last change. A short atom-count table, as above, catches all three. This list is my own experience-based advice (check).

What I did and did not read

The Chemical reaction page is 252,170 characters long and I read the first 100,000, which stops in the organic chemistry section; the later sections were not read. The Chemical equation and Conservation of mass pages were reported to me as complete. The Combustion page is 117,368 characters and I read the first 100,000. All quotations came

through the fetch tool's summary of each page, so check them against the pages before relying on them. Atom counts were run in Python. Statements marked "(check)" are my own knowledge or reasoning.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles Chemical reaction, Chemical equation, Conservation of mass and Combustion are the sources used here.
- **A school textbook.** Look for graded balancing practice and the idea of ionic equations.
- **Practical.** Write an equation for something you have seen happen at home, then do the atom-count table. If the counts do not match, the equation is wrong, not the world.

Exercises

EXERCISE 1

Balance and prove it. Balance $N_2 + H_2 \rightarrow NH_3$ (nitrogen and hydrogen forming ammonia), showing each step. Then make an atom-count table with the left and right totals for each element before and after balancing, and say what you would be doing wrong if you changed NH_3 to NH_2.

→ Solution

EXERCISE 2

Classify the reaction. For each balanced equation, name the type from this chapter (synthesis, decomposition, single displacement, double displacement or combustion) and give the one phrase from the type's description that justifies it: (a) $2H_2O_2 \rightarrow 2H_2O + O_2$; (b) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$; (c) $2HCl + 2Na \rightarrow 2NaCl + H_2$; (d) $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$. Say whether any fits more than one label.

→ Solution

EXERCISE 3

From description to equation. Magnesium, a solid metal, burns in oxygen gas to form the white solid magnesium oxide. Write a word equation, then a symbol equation with state symbols, using the Chapter 4 idea of ion charges to get the formula of magnesium oxide. Balance it, show an atom-count table, and name one thing the equation does not tell you about the burning.

→ Solution

Quick Reference

- **Reactants → products:** starting substances on the left, new substances on the right
- **Word equation / symbol equation:** words first, then formulae with coefficients
- **State symbols:** (s) solid, (l) liquid, (g) gas, (aq) aqueous solution
- **Balance:** same number of atoms of each element on both sides; change coefficients only, never subscripts
- **Why:** conservation of mass, confirmed by Lavoisier in the late 18th century, per the Conservation of mass article
- **Types:** synthesis, decomposition, single displacement, double displacement, combustion
- **Not shown by an equation:** whether it happens, how fast, conditions, and masses
- **Always:** recount every element after the last change

Coming next

Chapter 7 turns the counts in an equation into quantities you can weigh and measure: the mole, the Avogadro constant, molar mass and simple stoichiometry.

Chapter 6 ended with a limit: a coefficient in a balanced equation counts molecules, not grams. Yet in a laboratory you cannot count molecules; you can only weigh things. This chapter builds the bridge. It introduces the mole, the Avogadro constant and molar mass, then uses them in two worked examples: converting between grams and moles, and finding the mass of a product from a balanced equation. Each fact is attributed to the Wikipedia page that states it, and my own arithmetic or general knowledge is marked "(check)".

Why Count in Moles?

Chapter 2 gave atoms masses of a few dalton each, far too small to see or weigh one at a time (the scale is from Chapter 2's proton and neutron figures). Any sample you can pick up holds a vast number of them. So chemists use a counting unit big enough to match that scale, in the way a shop counts eggs by the dozen. The Mole (unit) article says the mole measures "the amount of substance", and defines the ratio $n = N/N_A$, the number of entities N divided by the Avogadro constant N_A, as "a measure of the amount of substance (with the unit mole)". The article says the "elementary entity" counted may be "an atom, a molecule, an ion, an ion pair, or a subatomic particle", so a mole of oxygen could mean a mole of atoms or a mole of O_2 molecules, and you must say which (that last warning is my own advice (check)).

The Avogadro Constant

After the 2019 changes, the Mole (unit) article says, one mole contains "exactly 6.02214076×10^{23} elementary entities". The same page says the constant's numerical value "has units of mol^{-1}", so the Avogadro constant is $6.02214076 \times 10^{23}$ mol^{-1}. In this course I round it to 6.02×10^{23} when the arithmetic allows (my rounding (check)). The same article says the name "mole" was "coined by the chemist Wilhelm Ostwald in 1894 from the German word Molekül". I did not fetch the Amedeo Avogadro article, so I make no claims here about the man himself.

The mole

Avogadro constant

The unit of amount of substance, symbol mol. One mole is $6.02214076 \times 10^{23}$ elementary entities (Mole (unit) article).

$N_A = 6.02214076 \times 10^{23}$ mol^{-1}. Number of particles $N = n \times N_A$ (the rearrangement of the page's $n = N/N_A$).

Molar mass

The Molar mass article defines it as "the ratio between the mass (m) and the amount of substance (n, measured in moles)": $M = m/n$, usually in g/mol.

The Molar mass article says that "the coherent unit of molar mass is kg/mol" but that "molar masses are almost always expressed with the unit g/mol". Use g/mol, as the worked examples below do.

Molar Mass and Relative Atomic Mass

This is where Chapter 2's relative atomic mass pays off. The Molar mass article says that for an element the molar mass "is given by the relative atomic mass $A_r(X)$ of the element multiplied by the molar mass constant", and that this constant is about 1 g/mol. In practice: the number is the same, only the unit differs. Chlorine's relative atomic mass is about 35.5 (Chapter 2), so a mole of chlorine atoms has a mass of about 35.5 g (my reading of the page's rule (check)).

For a compound, the page says the molar mass is "given by the sum of the relative atomic masses" of its elements, each multiplied by its atom count, then multiplied by the molar mass constant. Chapter 4 and Chapter 5 gave you the formulae to sum over. The page's own sodium chloride example is:

Copy

```
M(NaCl) = [22.98976928 + 35.453] x M_u
M(NaCl) = 58.443 g/mol      (Molar mass article: 58.443(2) g/mol; my Python sum: 58.44277)
```

The same article gives water's molar mass as 18.0153(3) g/mol. It shows no working for water, so I do not claim to reproduce its figures.

Worked Example: Grams to Moles and Back

The two conversions are one formula, n = m / M, rearranged. This example uses the Molar mass article's value for sodium chloride, 58.443 g/mol. All results were calculated in Python.

Question	Formula	Working	Answer
How many moles in 11.69 g of NaCl?	n = m / M	11.69 g ÷ 58.443 g/mol	0.2000 mol
What mass is 0.500 mol of NaCl?	m = n × M	0.500 mol × 58.443 g/mol	29.2 g
How many formula units in 0.500 mol?	$N = n \times N_A$	$0.500 \times 6.02214076 \times 10^{23}$	3.01×10^{23}

The units cancel as a check: g ÷ (g/mol) leaves mol. If your units do not cancel to the unit you wanted, you have the formula upside down (my advice (check)). Sodium chloride is ionic (Chapter 4), so the particles counted in the last row are formula units, not molecules; the choice of word is mine (check).

Worked Example: Mass of Product from a Balanced Equation

Chapter 6 balanced $2H_2 + O_2 \rightarrow 2H_2O$. Its coefficients now give mole ratios: the Stoichiometry article says "The molar ratio allows for conversion between moles of one substance and moles of another", with the coefficients supplying the ratio. The article's own example is copper and silver nitrate, which I follow in pattern but not in substance: it converts grams to moles, applies the ratio, then converts moles to grams. Question: how much water forms from 4.00 g of hydrogen, with oxygen in plenty?

Molar masses: for H_2 I use 1.00794 for hydrogen, the figure the Molar mass article uses in its sucrose example, so $M(H_2)$ = 2.01588 g/mol (my Python sum). For water I use that article's 18.0153 g/mol.

Step	What you do	Working	Result
1. Balanced equation	Check it first (Chapter 6)	$2H_2 + O_2 \rightarrow 2H_2O$	ratio $H_2 : H_2O = 2 : 2$
2. Grams to moles of H_2	n = m / M	4.00 g ÷ 2.01588 g/mol	1.984 mol

Step	What you do	Working	Result
3. Apply the mole ratio	2 mol H_2O per 2 mol H_2	1.984 mol × (2 / 2)	1.984 mol H_2O
4. Moles to grams of H_2O	m = n × M	1.984 mol × 18.0153 g/mol	35.7 g

So 4.00 g of hydrogen yields about 35.7 g of water, provided there is enough oxygen (Python gave 35.7468 g before rounding to three figures). The product is much heavier than the hydrogen because oxygen atoms join in, which is Chapter 6's conservation of mass at work (my reading (check)). The mole ratio here happens to be 1 : 1; in the Stoichiometry article's copper example it is 1 : 2 and the middle step matters more.

Common slips

Mixing up units, for example using kg or mg with a g/mol molar mass; using the formula mass of the wrong species, such as the atomic mass of oxygen (about 16) when the substance is O_2 (about 32); forgetting to balance the equation first, so the mole ratio is wrong; and dividing when you should multiply in step 2 or 4. Also check which entity you are counting (atoms, molecules or formula units). This list is my own experience-based advice (check).

What I did and did not read

I fetched the Mole (unit), Molar mass and Stoichiometry articles. The fetch tool reported none of them truncated and I cannot tell the full lengths; Mole (unit) gave no total length. Quotations came through the fetch tool's summary of each page, so check them against the pages. I did not fetch the Avogadro constant, Amedeo Avogadro, Relative atomic mass or 2019 SI revision pages for this chapter, so the 2019 definition is only as the Mole (unit) article reports it. Calculations were run in Python. Statements marked "(check)" are my own knowledge or reasoning.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles Mole (unit), Molar mass and Stoichiometry are the sources used here.
- **A school textbook.** Look for graded mole calculations and the idea of limiting reactants, which this chapter does not cover (the Stoichiometry example only says silver nitrate was in excess).
- **Practical.** Weigh a teaspoon of table salt on a kitchen scale and turn the mass into moles with the method above (the weighing is my suggestion (check)).

Exercises

EXERCISE 1

Grams to moles. Use C = 12.01 and O = 16.00 g/mol. Find the molar mass of carbon dioxide, CO_2, and then the number of moles in 22.0 g of it. Show the formula, the substitution and the unit cancelling.

→ Solution

EXERCISE 2

Moles to particles. Take 0.250 mol of water, H_2O, and $N_A = 6.02 \times 10^{23}$ mol^{-1}. How many molecules is that? How many atoms in total, and how many of those are hydrogen atoms? Give each answer in standard form.

→ Solution

EXERCISE 3

Mass of product. Magnesium burns in oxygen: $2Mg + O_2 \rightarrow 2MgO$. Use Mg = 24.31 and O = 16.00 g/mol. How many grams of magnesium oxide form from 6.00 g of magnesium, with oxygen in plenty? Set out the four steps used in this chapter.

→ Solution

Quick Reference

- **Mole:** the unit of amount of substance; 1 mol = $6.02214076 \times 10^{23}$ elementary entities (Mole (unit) article)
- **Avogadro constant:** $N_A = 6.02214076 \times 10^{23}\ \text{mol}^{-1}$
- **Molar mass:** $M = m / n$, in g/mol; for an element, the number equals the relative atomic mass
- **Compound:** add the molar masses of all the atoms in the formula
- **Grams to moles:** $n = m / M$. **Moles to grams:** $m = n \times M$
- **Particles:** $N = n \times N_A$
- **Mass from an equation:** balance, grams to moles, mole ratio from coefficients, moles to grams
- **Slips:** units, wrong formula mass, unbalanced equation

Coming next

Chapter 8 turns to acids, bases and the pH scale, where the amounts of substance from this chapter reappear when acids and bases are neutralised.

Chemistry Fundamentals

Course 1 · Chapter 8 · Acids, Bases and pH: Sour, Slippery and Logarithmic

Chapter 7 promised that amounts of substance would reappear when acids and bases are neutralised, and this chapter delivers that. It defines acids and bases in two simple ways, explains what the pH number measures and why a change of one unit is a tenfold change, tabulates the pH of everyday substances, and treats neutralisation as a reaction you can balance (Chapter 6) and that makes a salt (Chapter 4). Each fact is attributed to the Wikipedia page that states it, and my own arithmetic or general knowledge is marked "(check)".

Acids and Bases: Two Simple Definitions

The Acid article gives two definitions. The older, Arrhenius, says: "An Arrhenius acid is a substance that, when added to water, increases the concentration of H^+ ions in the water." The broader Brønsted–Lowry definition describes an acid as "a species that donates a proton to a Brønsted–Lowry base." The Base (chemistry) article mirrors these. An Arrhenius base is, in its words, "a substance which dissociates in aqueous solution to form hydroxide ions OH^-" (credited there to Arrhenius, 1884), and a Brønsted–Lowry base "is a substance that can accept hydrogen cations (H^+)" (1923). That page adds that this is the broader definition and that Arrhenius bases are a subset of it.

A hydrogen atom that has lost its one electron is a bare proton (Chapter 2's electron and proton counts), so "donates H^+" and "donates a proton" mean the same thing here (my link to earlier chapters (check)). The Hydronium article says that the ion produced when water is protonated is H_3O^+, but also that the aqueous proton is more accurately represented as $H^+(aq)$, and it suggests using the symbol $H^+(aq)$ where possible. This course writes H^+ for short, as the pH formula pages do (my choice (check)).

Arrhenius acid

Adds H^+ ions to water (Acid article).

Arrhenius base

Forms hydroxide ions, OH^-, in water (Base (chemistry) article).

Brønsted–Lowry

> An acid donates a proton; a base accepts one (Acid and Base (chemistry) articles).

On everyday properties, the Acid article says acids "form aqueous solutions with a sour taste" and "can turn blue litmus red". The Base (chemistry) article says bases are "slippery to the touch, can taste bitter", and "turn red litmus paper blue". Tasting is not a test you should use, as the safety box below explains.

Strong and Weak

I take only what the pages say. The Acid article: "A strong acid is one that completely dissociates in water", while a weak acid "only partially dissociates", so both the acid and its conjugate base are present at equilibrium. The Base (chemistry) article: "A strong base is a base that is quantitatively protonated upon exposure to water", while a weak base only partly accepts protons. That article notes that "strong" here names a category and does not describe a qualitative intensity. So strong and weak describe how fully a substance reacts with water, not how concentrated the solution is (my restatement of that point (check)). "Equilibrium" is a topic for Chapter 9.

What pH Measures

The PH article defines pH as "the decimal logarithm of the reciprocal of the hydrogen ion activity, a_{H+}", that is, $pH = -\log_{10}(a_{H+})$. It says the range is "commonly given as 0 to 14". The Hydronium article gives a simpler working form, $pH = -\log([H_3O^+]/M)$, where M is mol/L. That page also says that the molar concentration of hydronium or H^+ ions determines a solution's pH. "Activity" is a refinement of concentration that I do not cover (check); this chapter uses concentration in mol/L, a unit built on Chapter 7's mole (my simplification (check)).

The key idea is the logarithm. The PH article states that a difference of one pH unit is equivalent to a tenfold difference in hydrogen ion concentration. The lower the pH, the higher the concentration. Here are some worked numbers, with every pH calculated in Python as $-\log_{10}$ of the concentration (my arithmetic (check)):

$[H^+]$ in mol/L	As a power of ten	pH	Compared with the row above
0.1	10^{-1}	1	(start)
0.001	10^{-3}	3	100 times fewer ions (two units up)
0.0000001	10^{-7}	7	10 000 times fewer ions (four units up)
0.00032	3.2×10^{-4}	3.49	between pH 3 and pH 4 (3.4949 before rounding)

Pure water is the middle of the scale: the PH article says "Pure water has a pH of 7 at 25 °C, meaning it is neutral." In the table that is a concentration of 10^{-7} mol/L. So pH 3 versus pH 6 is a factor of $10^3 = 1000$ in hydrogen-ion concentration (Python: 1000.0), not "twice as acidic" or "half as much". Values below 7 are acidic, above 7 basic (the neutral-point pairing follows from the page's table, which labels pure water as neutral and ammonia as a base).

The pH of Common Substances

The PH article lists these "average pH of common solutions", reproduced as the fetch tool returned them. Ranges are given as the page gives them, because real samples vary.

Substance	pH (PH article)	Labelled on the page
Battery acid	< 1	Acid
Gastric acid	1.0–1.5	
Vinegar	2–3	
Orange juice	3.3–4.2	
Black coffee	5–5.03	
Milk	6.5–6.8	
Pure water at 25 °C	7	Neutral
Sea water	7.5–8.4	

Substance	pH (PH article)	Labelled on the page
Ammonia	11.0–11.5	Base
Bleach	12.5	
Lye	14	

Using the logarithm rule, take gastric acid at the top of its listed range (pH 1.5) and orange juice at the top of its listed range (pH 4.2). The difference is 2.7 units, so the hydrogen-ion concentration of the gastric acid is about $10^{2.7}$ times larger, which Python gave as 501.19 (roughly 500 times). This comparison is my own calculation and the page's ranges are wide, so treat it as an order of magnitude (check).

Neutralisation: Acid + Base → Salt + Water

The Neutralization (chemistry) article defines it as "a chemical reaction in which acid and a base react with an equivalent quantity of each other". It says the reaction was historically written as "acid + base (alkali) → salt + water", and its example is HCl + NaOH → NaCl + H_2O. That equation is already balanced, as Chapter 6's atom count shows: one each of H, Cl, Na and O on both sides, with H counted as 2 on each side (my check of the page's equation (check)). By Chapter 6's classification it also has the double displacement pattern (my reading (check)).

The salt is sodium chloride, Na^+ and Cl^- in a lattice (Chapter 4). A "salt" in chemistry means an ionic compound of this kind, not only table salt (my general knowledge (check)). In the Brønsted–Lowry picture, the acid HCl donates a proton, H^+, and the hydroxide ion takes it to make water (my fit of the definitions to the equation (check)).

Copy

```
acid + base  →  salt + water
HCl + NaOH  →  NaCl + H2O      (Neutralization article's example)
H2SO4 + 2NaOH  →  Na2SO4 + 2H2O   (my own example, balanced by atom count (check))
```

Chapter 7's mole ratios apply directly. In HCl + NaOH → NaCl + H_2O the coefficients are 1 : 1, so 0.0250 mol of HCl needs 0.0250 mol of NaOH. With M(NaOH) of about 22.99 + 16.00 + 1.01 = 40.00 g/mol (my sum of rounded masses (check)), that is 0.0250 × 40.00 = 1.00 g of NaOH (Python). For the second equation above, the ratio is 1 : 2, so 0.0250 mol of H_2SO_4 needs 0.0500 mol of NaOH (Python: 0.0500).

The page also says how the final pH depends on strength. For a strong acid with a strong base, "the pH of such a solution is close to a value of 7"; with a weak acid and a strong base, the pH at the end-point "is greater than 7". So neutral pH is not guaranteed after "neutralising", and I do not explain the reason here (check).

Indicators and Litmus

The pH indicator article describes an indicator as "a halochromic chemical compound added in small amounts to a solution" that lets you determine pH visually or spectroscopically. Litmus, it says, is "a naturally occurring pH indicator made from a mixture of lichen species". The Litmus article adds that litmus is "a water-soluble mixture of different dyes extracted from lichens", often absorbed into filter paper, and that "in an acidic medium, blue litmus paper turns red" while "in a basic or alkaline medium, red litmus paper turns blue". Both pages give the colour change as occurring over pH 4.5 to 8.3 (the Litmus article: "at 25 °C"). The indicator article's other transitions are below. Universal indicator, it says, is used "when only rough knowledge of pH is necessary".

Indicator	Colours	pH range (pH indicator article)
Litmus	red to blue	4.5 to 8.3
Methyl orange	red to yellow	3.1 to 4.4
Phenolphthalein	colourless to purple-pink	8.3 to 10.0

Universal indicator, pH	Meaning on the page	Colour
1–3	strong acid	red
3–6	weak acid	orange/yellow
7	neutral	green
8–11	weak alkali	blue
11–14	strong alkali	violet/indigo

A practical consequence of the ranges (my reasoning (check)): litmus changes across a wide band, so it tells you "acid" or "base" but not a precise value, while an indicator that changes at pH 8.3 to 10.0 only reveals that a solution is clearly basic.

Safety

The Acid article says "strong acids and some concentrated weak acids are corrosive", and the Base (chemistry) article says "concentrated or strong bases are caustic on organic matter". It also says mixing strong bases with acid spills "can cause a violent exothermic reaction". The following is general advice from me (check): never taste or smell-test a chemical to identify it, wear eye protection when handling any acid or base, and never mix household cleaning products, because the combinations can be dangerous. The Acid article also notes that rhubarb leaves and unripe carambolas are toxic because of their oxalic acid, so "sour" does not mean "safe".

What I did and did not read

I fetched the PH, Acid, Base (chemistry), Neutralization (chemistry), PH indicator, Litmus and Hydronium articles. The PH article's text was 119 881 characters and I read only the first 100 000, which ended in its reference list, so the last 19 881 were not read. The Acid article was 103 226 characters and I read the first 100 000; the final 3 226 were not read. The fetch tool reported the other pages as complete, but I cannot tell their lengths. Quotations came through the fetch tool's summary of each page, so check them against the pages. Calculations were run in Python. Statements marked "(check)" are my own knowledge or reasoning.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles above are the sources used here; the PH article also covers measurement and the activity idea this chapter skips.
- **A school textbook.** Look for titration, buffers and the acid dissociation constant, none of which are covered here.
- **Practical.** Ask a supervised science teacher or use a kit with litmus paper to test safe, dilute items such as those in the table; do not improvise with cleaning products (my suggestion (check)).

Exercises

EXERCISE 1

pH from a concentration. Use pH = $-\log_{10}[H^+]$ with $[H^+]$ in mol/L. Find the pH of (a) a solution with $[H^+] = 3.2 \times 10^{-4}$ mol/L and (b) a solution with $[H^+] = 1.0 \times 10^{-9}$ mol/L. Say whether each is acidic or basic, and compare each with pure water at pH 7.

→ Solution

EXERCISE 2

Compare by a factor. Using the rule that one pH unit is a factor of ten: (a) how many times higher is the hydrogen-ion concentration at pH 1 than at pH 4? (b) How many times higher at pH 3 (vinegar) than at pH 5 (black coffee)? (c) A solution goes from pH 3 to pH 6. By what factor does the hydrogen-ion concentration change, and in which direction?

→ Solution

EXERCISE 3

Neutralisation equation. Calcium hydroxide, $Ca(OH)_2$, reacts with hydrochloric acid, HCl, to give a salt and water. Write the unbalanced equation, balance it with Chapter 6's atom count, name the salt using the ions of Chapter 4 (Ca^{2+} and Cl^-), and state how many moles of HCl react with 0.010 mol of $Ca(OH)_2$.

→ Solution

Quick Reference

- **Arrhenius:** an acid adds H^+ to water; a base gives OH^-
- **Brønsted–Lowry:** an acid donates a proton; a base accepts one

- **pH:** $-\log_{10}$ of the hydrogen-ion concentration; range commonly 0 to 14; pure water 7 at 25 °C
- **One pH unit:** a tenfold change in hydrogen-ion concentration
- **Strong or weak:** how completely the substance reacts with water, not how concentrated it is
- **Neutralisation:** acid + base → salt + water; balance it, then use mole ratios
- **Litmus:** red in acid, blue in base; changes over pH 4.5 to 8.3
- **Safety:** strong acids and bases are hazardous; never taste or mix chemicals

Coming next

Chapter 9 turns to energy and rates: why some reactions release heat and others absorb it, what activation energy and catalysts do, and what equilibrium means (the word used above for weak acids).

Chapter 6 said that a balanced equation does not tell you how fast a reaction goes or whether it happens at all, and that most reactions can run in both directions. This chapter delivers on that. It separates three different questions: does a reaction give out or take in energy, how fast does it go, and does it run both ways until a balance is reached. Each fact is attributed to the Wikipedia page that states it, and my own arithmetic or general knowledge is marked "(check)". Any energy numbers in the exercises are invented teaching values, not real data.

Exothermic and Endothermic

The Exothermic process article defines an exothermic process as a "thermodynamic process that releases energy to its surroundings", usually as heat, though it says light, electricity or sound are also possible. It also describes the IUPAC wording: a reaction whose overall standard enthalpy change is negative. The Endothermic process article gives the opposite: "An endothermic process is a chemical or physical process that absorbs heat from its surroundings", and says an endothermic reaction "generally leads to a decrease in the temperature of the surroundings". So, from the point of view of the surroundings (the flask, the water, your hand), exothermic means it warms up and endothermic means it cools down (my restatement of those two pages (check)).

Exothermic

Energy leaves the reacting system for the surroundings; the enthalpy change is negative (Exothermic process article).

Endothermic

Heat is absorbed from the surroundings, which tend to get colder (Endothermic process article).

Physical or chemical

Both pages include physical changes as well as reactions, tying back to Chapter 1's physical and chemical change.

Type	Examples the pages list
Exothermic (Exothermic process article)	burning fuels such as wood, coal and oil; the thermite reaction; alkali metals reacting with water; rain condensing from water vapour; mixing water with strong acids or bases; acid-base reactions; cement and concrete setting; zinc reacting with hydrochloric acid; cellular respiration
Endothermic (Endothermic process article)	evaporation; sublimation; cracking of alkanes; thermal decomposition; hydrolysis; dissolving ammonium nitrate in water; the melting of ice cubes; dissolving barium hydroxide and ammonium chloride together

Notice that acid-base reactions are on the exothermic list. That fits Chapter 8, where the safety note said mixing strong bases with acid spills can cause a violent exothermic reaction. Combustion, from Chapter 6, is the everyday exothermic example, and the Combustion wording used there ("often producing energy in the form of heat or light") says the same thing. Melting and evaporation are endothermic, which is why a change of state (Chapter 1) needs heat to be supplied (my link between the pages and Chapter 1 (check)).

Bonds: Energy to Break, Energy Released When Made

Chapter 5 described covalent bonds as shared electrons. The Bond-dissociation energy article defines bond dissociation energy as "the standard enthalpy change when A−B is cleaved by homolysis to give fragments A and B". The page gives example values with their units: for H−H, 435.780 kJ/mol at 298 K in the main text and 431 kJ/mol in its table (the page gives both, and I did not resolve the difference); for O=O, 498 kJ/mol in its table. That page does not state in so many words that breaking bonds takes in energy; the values are positive enthalpy changes for splitting a bond, which implies it (the page's own limitation). The Collision theory article does say that colliding particles must have enough energy "to break existing bonds and form new ones".

The usual textbook picture is that a reaction breaks bonds in the reactants, which takes in energy, and makes bonds in the products, which gives energy out, and the overall result depends on which is larger. The bond-energy page does not state that sum as a rule in the part I read, so treat it as general knowledge (check). The per-mole unit, kJ/mol, builds on Chapter 7's mole (check). The sign convention used in this chapter is that a negative energy change means energy leaves the system, matching the negative enthalpy change the Exothermic process article describes.

Picture

Think of a balance: energy in to break the old bonds, energy out when the new bonds form. Net out means exothermic, net in means endothermic. This is a teaching summary, and the exact bookkeeping (enthalpy, average bond energies) is more careful than this (check).

Activation Energy and an Energy Profile

A reaction can give out energy overall and still not start by itself, because the reactants must first get over a barrier. The Activation energy article defines it as "the minimum energy barrier that reactant molecules must overcome to transform into products". It describes an energy profile that plots energy against reaction coordinate, in which "the highest energy position (peak position) represents the transition state". The next table is such a profile written as text. Every number is an invented teaching value, not a measured one.

Stage of the reaction	Energy level (invented, kJ/mol)	What it means
Reactants	0	starting level
Transition state (the peak)	+60	top of the barrier
Products	−250	lower than the reactants, so exothermic

From these invented levels, the activation energy is 60 − 0 = 60 kJ/mol, and the overall energy change is −250 − 0 = −250 kJ/mol (Python). The two numbers answer different questions: 60 is how high the barrier is, and −250 is where you end up. The reverse reaction, products back to reactants, would face a barrier of 60 − (−250) = 310 kJ/mol (Python), which is my own reading of the picture (check).

Do not confuse the two numbers

A large energy release does not mean a fast reaction, and a small barrier does not mean a large release. Energy change is about the start and end levels. Rate (below) is mostly about

the barrier and how often particles meet. The sources treat these as separate ideas; the Chemical reaction article quoted in Chapter 6 puts speed under reaction kinetics and "whether it happens" under free energy, and I do not go further into those (check).

What Changes the Rate

The Reaction rate article lists the factors, and I quote what its summary of the page gave. The page's length was not stated.

Factor	What the Reaction rate article says
Concentration	"Reaction rate increases with concentration, as described by the rate law and explained by collision theory."
Temperature	A higher temperature increases the rate; the rule of thumb is that "a rise of ten degrees Celsius results in approximately twice the reaction rate constant."
Surface area	"The rate of reaction increases as the surface area does": more solid particles are exposed to reactant molecules.
Catalyst	"The presence of a catalyst increases the reaction rate", by offering a pathway with lower activation energy.
Pressure	"The rate of gaseous reactions increases with pressure"; for condensed phases the dependence is weak.
Light	More intense light means particles absorb more energy, so the reaction goes faster.

Applying the page's rule of thumb: a rise of 20 °C is two steps of 10 °C, so the rate constant roughly doubles twice, a factor of $2^2 = 4$ (Python: 4.0). A rise of 30 °C gives $2^3 = 8$ (Python: 8). This is a rough rule, as the page's word "approximately" says, and it is not exact for every reaction (check).

Collision Theory in Plain Terms

The Collision theory article says that for a reaction, "suitable particles of the reactant hit each other with the correct orientation", and "the successful collisions must have enough energy, also

known as activation energy". It also says increasing concentration "brings about more collisions and hence more successful collisions", and that increasing temperature "increases the average kinetic energy of the molecules in a solution". Put together (my plain-language summary (check)):

- Particles must meet, so more crowded particles (higher concentration, or more exposed surface) mean more meetings.
- They must meet the right way round, with enough energy to get over the barrier.
- Heating gives more particles enough energy to react, so more of the collisions succeed.

Most collisions do nothing. The page says only some of the impacts produce a noticeable change.

Catalysts

The Catalysis article defines catalysis as "the increase in rate of a chemical reaction due to an added substance known as a catalyst", and says "Catalysts are not consumed by the reaction and remain unchanged after the reaction." It says the catalyst "provides an alternative reaction mechanism (reaction pathway) having a lower activation energy". The Activation energy article adds that a catalyst "does not change the energies of the original reactants or products", so it lowers the peak of the profile but not the start and end levels. The Catalysis article's examples are enzymes ("protein-based catalysts in metabolism and catabolism") and catalytic converters, "typically composed of platinum and rhodium", which "break down some of the more harmful byproducts of automobile exhaust".

Using the invented profile from earlier, a catalyst that lowered the peak from +60 to +40 would cut the forward barrier from 60 to 40 kJ/mol, and leave the overall change at −250 kJ/mol (Python). It would give no extra energy out. That is an invented example of what the pages describe, not a measured one. Because the catalyst is not used up, it does not appear as a reactant or product in the balanced equation of Chapter 6; the Chemical equation article, as Chapter 6 noted, shows catalysts only above or below the arrow.

Reversible Reactions and Equilibrium: A First Look

Chapter 6 reported that most reactions "can and do run in both directions". The Chemical equilibrium article describes the state in which reactants and products are both present, with no

net change in their concentrations: "This state results when the forward reaction proceeds at the same rate as the reverse reaction", known as dynamic equilibrium. It says arrows pointing both ways in an equation indicate equilibrium, and credits Jacobus van 't Hoff with proposing the double arrow, ⇌, in 1884. The word "dynamic" matters: both reactions are still going on, only their rates match (my plain restatement (check)).

Chapter 8's weak acids are the example the end of that chapter pointed to: a weak acid only partly dissociates, so both the acid and its conjugate base are present at equilibrium (Acid article, as quoted in Chapter 8).

The Le Chatelier's principle article says that if the equilibrium of a system is disturbed by a change in one or more of the determining factors, the system adjusts to counter the change. The words I could see on the page were only part of the statement, so this is a simplified version (check). From the page's summary:

- **Concentration.** Adding carbon monoxide to the methanol synthesis, $CO + 2\ H_2 \rightleftharpoons CH_3OH$, pushes the reaction toward more methanol, and removing a product can drive the reaction forward.
- **Temperature.** The direction depends on whether the reaction is exothermic or endothermic. The page says for an exothermic reaction "an increase in temperature decreases the equilibrium constant".
- **Example.** The reaction $2NO_2 \rightleftharpoons N_2O_4$ is exothermic. Cooling the flask shifts it toward N_2O_4 and the flask becomes colourless; heating it shifts the reaction back toward NO_2 and the flask turns reddish-brown.
- **Pressure.** The page says the equilibrium concentrations "do not directly depend on the total pressure"; pressure matters mainly through partial pressures, and has no effect when the gas moles are equal on both sides.

Equilibrium constants and calculations are beyond this course. Equilibrium also is not "equal amounts": the page says only that concentrations stop changing (my reading (check)).

> **Caution: rate is not energy change**
>
> Three questions, three different answers. Is it exothermic or endothermic (the energy change)? How fast is it (the rate, set by concentration, temperature, surface area, catalyst and the activation energy)? Does it reach a balance (equilibrium)? A catalyst changes the rate and the barrier, not the overall energy change. Raising the temperature speeds nearly

any reaction, but for an equilibrium it can also change which side is favoured (per the Le Chatelier's principle article).

What I did and did not read

I fetched the Exothermic process, Endothermic process, Activation energy, Reaction rate, Collision theory, Catalysis, Chemical equilibrium, Le Chatelier's principle and Bond-dissociation energy articles. The Catalysis article's text was 130 382 characters and I read the first 100 000; the final 30 382 were not read. The Chemical equilibrium article was 107 274 characters and I read the first 100 000; the final 7 274 were not read. For the other seven pages the fetch tool did not state a length, so I cannot tell how much of each I read. Quotations came through the fetch tool's summary of each page, so check them against the pages. Calculations were run in Python. Statements marked "(check)" are my own knowledge or reasoning.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles above are the sources used here; the Activation energy article covers the Arrhenius relation, which this chapter leaves out.
- **A school textbook.** Look for enthalpy changes, energy-profile diagrams, rate laws and equilibrium constants.
- **Practical.** Only with supervision and a proper kit: compare how fast a safe, dilute reaction goes warm and cold (my suggestion (check)).

Exercises

EXERCISE 1

Classify by what is observed. Say whether each change is exothermic or endothermic, using the definitions on this page: (a) a flask warms up from 20 °C to 35 °C when two solutions are mixed; (b) a pool of water on a table evaporates; (c) a bottle of dissolved ammonium nitrate and water gets colder to the touch; (d) wood burns and gives out light and heat; (e) rain forms from water vapour. For each, say whether the surroundings gain or lose heat.

→ Solution

EXERCISE 2

Energy bookkeeping with invented numbers. Reaction X takes in 1000 kJ to break the bonds of the reactants and gives out 1250 kJ when the new bonds form. Reaction Y takes in 900 kJ and gives out 700 kJ. For each, work out the overall energy change (energy in minus energy out, so negative means energy leaves) and say if it is exothermic or endothermic. Then, using an invented profile for X with reactants at 0, the peak at +60 and products at the level you found, give the activation energy.

→ Solution

EXERCISE 3

Speeding a reaction. A reaction between a lump of a solid and a solution is too slow. Four changes are proposed: (1) crush the solid to a powder; (2) cool the mixture by 10 °C; (3) add a suitable catalyst; (4) use a more dilute solution. Say which would speed it up and which would slow it, and give the collision-theory reason for each. Then, using the rule of thumb of this chapter, estimate by what factor the rate constant changes if the temperature is raised by 20 °C instead.

→ Solution

Quick Reference

- **Exothermic:** releases energy to the surroundings; negative enthalpy change
- **Endothermic:** absorbs heat from the surroundings, which usually cool
- **Activation energy:** the minimum energy barrier to turn reactants into products
- **Rate goes up with:** concentration, temperature, surface area, a catalyst (and pressure for gases, light)
- **Rule of thumb:** about double the rate constant per 10 °C rise (approximate)
- **Collision theory:** particles must collide, the right way round, with enough energy

- **Catalyst:** speeds the reaction, is not consumed, lowers the barrier, leaves the overall energy change alone
- **Equilibrium:** forward and reverse rates equal, concentrations steady; written with ⇌
- **Le Chatelier:** a disturbed equilibrium adjusts to counter the change

Coming next

Chapter 10 is the capstone: it uses everything so far on cooking, cleaning, materials, reading a label and weighing everyday chemical claims.

Chemistry Fundamentals

Course 1 · Chapter 10 · Capstone: Chemistry in Everyday Life

Nine chapters of ideas, and this one puts them to work on things you already own: a tub of baking soda, a bar of soap, a rusty bike chain, a bottle of bleach, the label on a packet. The aim is not new theory but a habit: for any everyday chemical claim, name which earlier chapter explains it, and check what a source actually says. Each fact below is attributed to the Wikipedia page that states it. My own arithmetic or general knowledge is marked "(check)", and the numbers in the exercises are invented teaching values.

Example 1: Baking Soda and an Acid

The Sodium bicarbonate article gives the reaction as "NaHCO3 + H+ → Na+ + CO2 + H2O", and says "When it reacts with acid or is heated, carbon dioxide is released." In baking, it says, baking soda "is primarily used in baking as a leavening agent": the gas makes batters expand. Now tie it back:

- **Chapter 8:** the H^+ on the left is the acid. Baking soda is a base reacting with an acid, so a sour ingredient in a recipe (vinegar, lemon juice, buttermilk) supplies it (my link to Chapter 8 (check)).
- **Chapter 6:** the equation must balance. Count the charge and atoms: one sodium, one hydrogen from the soda plus one H^+ gives H_2O and CO_2, with matching totals each side (my check of the page's equation (check)).
- **Chapter 7:** using relative atomic masses (Python), the molar mass of $NaHCO_3$ is 84.01 g/mol and of CO_2 is 44.01 g/mol. One mole of soda gives one mole of CO_2, so 10 g of baking soda (0.119 mol) can give at most 5.24 g of CO_2 if there is enough acid (Python).

The same page also says baking soda can be used as "a gentle scouring powder" when made into a paste with minimal water, and that for odour control in refrigerators "there is little evidence that it is effective in this application". That second sentence is the course's checking habit in miniature: a popular tip, and a source saying the evidence is weak.

Example 2: Cooking and Browning

The Maillard reaction article defines the reaction as "a chemical reaction between amino acids and reducing sugars to create melanoidins", and says it is "responsible for many colors and flavors in foods", for example seared meat, bread crusts, fried onions and roasted coffee. It says the reaction "typically proceeds rapidly from around 140 to 165 °C (280 to 330 °F)", and that "Caramelization is an entirely different process from Maillard browning": caramelization involves sugars alone, while the Maillard reaction needs amino acids too.

Links: this is a chemical change, not a physical one (Chapter 1), because new substances with new colours and smells form. The temperature range is the Chapter 9 idea that heating speeds reactions; many reactions need enough energy to get over a barrier (Chapter 9). Frying versus boiling then makes sense: boiling water stays near 100 °C (my general knowledge (check)), well below that range, which is why boiled food does not brown the same way (my inference, not stated by the page (check)).

Example 3: Cleaning with Soap

The Soap article says a soap molecule has a hydrophilic (water-attracting) end and a fat-attracting end, and that "The insoluble oil/fat 'dirt' becomes associated inside micelles", so grease can be carried away with water. It also says soap "emulsifies oils, enabling them to be carried away by running water." On making it: "An alkaline solution (often lye) induces saponification", in which "the triglyceride fats first hydrolyze into salts of fatty acids", with glycerol released. On hard water: "When used in hard water, soap does not lather well, but forms soap scum."

Chapter 5: polarity

Water is polar and grease is not, so they do not mix. A soap molecule has one end that suits water and one that suits fat, bridging the two (my link, consistent with the page's description (check)).

Chapter 4: ions

The page calls soaps "salts of fatty acids", so they are ionic compounds with a metal cation, tying to Chapter 4 (my link (check)).

Chapter 8: bases

Lye is the alkaline (base) ingredient in saponification, so making soap is a base

reacting with a fat (Chapter 8 vocabulary, my link (check)).

Example 4: Rust and Sea Water

The Rust article describes rust as "an iron oxide, a usually reddish-brown oxide formed by the reaction of iron and oxygen". It says "When iron is in contact with water and oxygen, it rusts", and calls it "an oxidation reaction specifically occurring with iron" in which iron gives up electrons to oxygen. The summary I read said salt speeds the process, so a bicycle left by the sea rusts faster than one kept inside (my everyday reading of that point (check)). For prevention it names galvanizing, "a layer of metallic zinc", and coatings "such as paint, lacquer, varnish", which keep air and water off the iron.

A simplified balanced equation is $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$: four iron and six oxygen atoms on each side (Python atom count). Real rust is a mix of hydrated oxides and oxide-hydroxides according to the page, so this equation is a teaching simplification (check). Links: it is a combination reaction type of Chapter 6; iron is a metal giving up electrons as in Chapter 4's ions; and rust's speed is a Chapter 9 question, since a paint coat or zinc layer lowers how often iron meets water and oxygen (my link (check)). Sea water is a mixture of water with dissolved salts (Chapter 1), and its sodium chloride is the ionic compound of Chapter 4 (formula mass 58.44 g/mol using 22.990 + 35.45, Python).

Reading an Ingredient List or a Hazard Label

I could not verify a general rule for how ingredient lists are ordered from the pages I read, so treat this as general knowledge (check): in many countries ingredients must be listed in descending order of weight, so the first item is the largest part. Use the list like this:

1. Find the **substance names**. Many are plain names of compounds from Chapters 4 and 5 (sodium chloride is table salt).
2. Remember a name is not a warning. A long chemical name is just a name; water is also "dihydrogen monoxide" and harmless (my example (check)).
3. Look for **percentages**. The Bleach article says household bleach is typically "Sodium hypochlorite, usually as a 3–6% solution in water". So in 500 g of a 5 per cent solution there is 500 × 0.05 = 25 g of the active compound (Python).

For hazard labels, the GHS hazard pictograms article describes "a standard set of hazard warning images" in the Globally Harmonized System, with nine pictograms, GHS01 to GHS09. I summarise it in the next table; the page said it is based on older editions, so the latest rules may differ (the page flags itself as outdated).

Pictogram	Meaning on the page
GHS01	Explosive
GHS02	Flammable (example: acetone)
GHS03	Oxidizing (example: hydrogen peroxide)
GHS04	Compressed gas
GHS05	Corrosive: corrodes metals or causes severe skin or eye damage
GHS06	Toxic: skull and crossbones, severe acute toxicity
GHS07	Exclamation mark: less severe acute toxicity, skin and eye irritation
GHS08	Serious health hazard (for example carcinogenicity, organ toxicity)
GHS09	Hazardous to the aquatic environment

The page also says labels carry the signal word "Danger" or "Warning" where necessary; it did not say which pictograms take which word. The corrosive pictogram links to Chapter 8's strong acids and bases (my link (check)). Also read the product's own instructions and safety sheet, which this course cannot replace.

Safety at home

Never mix household cleaners. The Bleach article says "Mixing an acid cleaner with a hypochlorite bleach can cause toxic chlorine gas to be released", that bleach "should never be mixed with vinegar or other acids", and that mixing bleach with ammonia produces nitrogen trichloride, described as potentially explosive and irritating, along with chloramines that are "very irritating to the eyes and lungs". Other general advice (my own, not from a page I read (check)): keep products in their original labelled containers, out of reach of children and pets, with the lid on; ventilate when cleaning; do not store them near food or heat; read the label before first use; and if someone is hurt or swallows something,

contact local emergency or poison advice straight away. This course is not a safety manual.

Judging Claims: "Chemical-Free" and Others

Chapter 1 defined chemistry as the study of matter, and all matter is made of substances. The Chemophobia article says "chemical free" is "impossible if taken literally, since all consumer products consist of chemical substances", and that such labels appeal to a fear by suggesting a safer alternative. Chemists quoted there say the divide between natural and artificial is "incorrect since all things in the universe are chemical by default", and a 1992 *Science* paper cited on the page states that "every natural chemical is also toxic at some dose". The page also describes IUPAC's definition of chemophobia as "irrational fear of chemicals".

That does not mean every claim is nonsense, nor that every synthetic product is safe. It means the word "chemical" alone tells you nothing. A better set of questions, in the spirit of the course's "check" habit:

Claim type	Question to ask
"Chemical-free"	Free of what, exactly? Water and air are chemicals too.
"Natural, so safe"	Natural substances can be harmful; the dose and the substance matter.
"Contains a chemical you can't pronounce"	What is the substance, in what amount, and what does a source say about it?
"Cures" or "detoxes"	Which study, how many people, compared with what?
"Tip from the internet"	Does a source test it? Compare the refrigerator baking soda claim above.
Big numbers without units	Per what amount? A percentage of what (see the bleach example)?

Hold the line

Ask three things of any chemical claim: which substance exactly, how much, and who says so. Then name the chapter that explains the idea, and whether a source tested it.

What I did and did not read

I fetched the Sodium bicarbonate, Soap, Maillard reaction, Rust, GHS hazard pictograms, Chemophobia and Bleach articles. The Sodium bicarbonate article's text was 153 981 characters and I read the first 100 000; the final 53 981 were not read. The Soap article was 102 275 characters and I read the first 100 000; the final 2 275 were not read. The Bleach article was 102 670 characters and I read the first 100 000; the final 2 670 were not read. For the other four pages the fetch tool did not state a length, so I cannot tell how much of each I read. Quotations came through the fetch tool's summary of each page, so check them against the pages. Calculations were run in Python. Statements marked "(check)" are my own knowledge or reasoning.

Where to Read More

- **Encyclopaedias.** The Wikipedia articles above are the sources used here.
- **A school or open-university chemistry course.** Look for organic chemistry (the chemistry of fats and soaps), electrochemistry (rusting) and food science (browning). These are general directions, not courses I checked.
- **Chapters of this course.** Return to Chapter 6 (equations) and Chapter 7 (the mole) for the calculation habits, and Chapter 8 for acids and bases.
- **Practical.** Only with supervision and safe materials: compare a wet and a dry nail over some days (my suggestion (check)).

Exercises

EXERCISE 1

Explain an everyday reaction using three chapters. A recipe adds 5.0 g of baking soda ($NaHCO_3$) to a batter containing lemon juice, and the batter rises. (a) Write the page's ionic equation for the reaction and say which chapter explains why a sour ingredient is needed. (b) Using molar masses of 84.01 g/mol for $NaHCO_3$ and 44.01 g/mol for CO_2,

find the largest mass of CO_2 it could give if the acid is not limiting. (c) Say whether this is a physical or a chemical change and why (Chapter 1).

→ Solution

EXERCISE 2

Interpret an invented label. A bottle of "Sparkle Kitchen Spray" lists: water, sodium hypochlorite 4%, sodium hydroxide 0.5%, fragrance. It shows the corrosive pictogram (GHS05) and the signal word "Danger". (a) How many grams of sodium hypochlorite are in 250 g of the product? (b) What does the corrosive pictogram mean on the page's description? (c) A friend wants to add vinegar "to make it stronger". What do you tell them, and which source supports it?

→ Solution

EXERCISE 3

Critique an invented advert. "PureLife Detox Salts: 100% chemical-free, made only from natural ingredients, so completely safe. Removes toxins in one night." (a) What is wrong with "chemical-free"? (b) What is wrong with "natural, so safe"? (c) Write three questions you would ask before believing "removes toxins", using the three-question habit of this chapter.

→ Solution

Quick Reference

- **Baking soda + acid:** $NaHCO_3 + H^+ \rightarrow Na^+ + CO_2 + H_2O$ (Sodium bicarbonate article)
- **Browning:** Maillard reaction needs amino acids and reducing sugars; about 140 to 165 °C
- **Soap:** one end likes water, one likes fat; micelles carry grease away
- **Rust:** iron + oxygen with water; paint and zinc coatings slow it
- **Labels:** nine GHS pictograms; read the signal word and the percentages

- **Never mix:** bleach with acids or with ammonia
- **"Chemical-free":** impossible if taken literally; ask which substance, how much, who says so

Course complete

You have finished Chemistry Fundamentals. In one line each: Chapter 1 sorted matter into states, substances and mixtures, physical and chemical change, and measured it with density and SI units. Chapter 2 broke the atom into protons, neutrons and electrons, with atomic number, isotopes and relative atomic mass. Chapter 3 laid out the periodic table as groups, periods and metal or non-metal families, with Mendeleev's story. Chapter 4 showed how electron transfer makes ions and ionic lattices, and how to write formulae from charges. Chapter 5 covered shared electrons, Lewis structures, shapes and polar bonds. Chapter 6 wrote reactions as balanced equations that conserve mass, in several types. Chapter 7 counted particles with the mole, molar mass and amounts from equations. Chapter 8 defined acids, bases, pH, neutralisation and indicators. Chapter 9 separated energy change, reaction rate and equilibrium. This chapter applied them to kitchens, cleaning, materials, labels and claims. Where next: re-read Chapters 6 and 7 until balancing and mole calculations feel routine, then try a school or open-university chemistry course in general terms (organic chemistry, electrochemistry and food science are natural follow-ons, my suggestion (check)), and keep the habit of asking "which substance, how much, who says so" about every chemical claim you meet.