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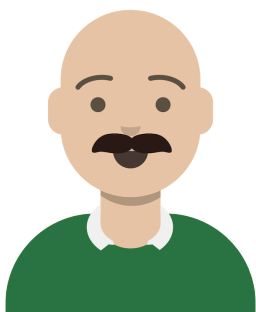
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Ready for Uni: An RMIT Chemistry Bridging Course

Ready for Uni: An RMIT Chemistry Bridging Course

*DURGA DHARMADANA; JACK BENCI;
MATTHEW MILLIS; JOSHUA MUIR; AND
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About this book

Purpose of this book

Chemistry is often known as the central science. Being the connecting bridge between biology, physics and nano-technology; chemistry plays a vital role in our daily and the operations of modern society. Chemists are expected to combine concepts and theories upon atoms and molecules that we can't even see! While a daunting task, it is one that can change the lives of billions. In this book, the authors aim to equip you with essential chemistry concepts needed to tackle tertiary studies, in this exploratory investigation into the world of chemical sciences.

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PART I

CHAPTER I: CHEMISTRY FOUNDATIONS

1.1 States of Matter

Learning Objectives

- Learn the basic terms used to describe matter.

What is Matter?

Matter encompasses anything that has mass and occupies space. This includes things that can be seen (all living and non-living, natural and synthetic things) and cannot be seen by the naked eye, like air and microorganisms. For instance, a book, a computer, food and dirt are all examples of matter. Sometimes, the matter may be difficult to identify. For example, air is matter, but because it is so thin compared to other matter (like books, computers, food, and dirt), we sometimes forget that air has mass and takes up space. In contrast, immaterial concepts and things like ideas, emotions and hopes do not qualify as matter.

Properties of Matter

To understand matter and how it changes, we need to be able to describe matter. There are two basic ways to describe matter:

physical properties and chemical properties. **Physical properties** are characteristics that describe matter as it exists, such as shape, colour, size and temperature. An important physical property is the phase (or state) of matter. The three fundamental phases of matter are **solid, liquid, and gas**. If we take water as an example, water can exist in all physical states: solid-ice cubes, liquid-liquid water, and gas-water vapour. Let's see how each of these states differ from one another.

Solid phase: the particles in a solid are packed tightly together, preventing the movement of particles and locking them into position (as shown in Figure 1.1.1), so that they cannot pass each other. Solids have a definite shape and a definite volume. Ice cubes are an example of a solid.



Figure 1.1.1 Particle arrangement inside a solid.

Liquid phase: The particles in a liquid are close together. However, the distance between the particles allows them to move past each other. Liquids take on the shape of the container, as they have an indefinite shape. Liquids have a definite volume. Think about a glass of water. The water takes the shape of the glass it is placed in, and we can define its volume, for instance, half a cup and a quarter cup.

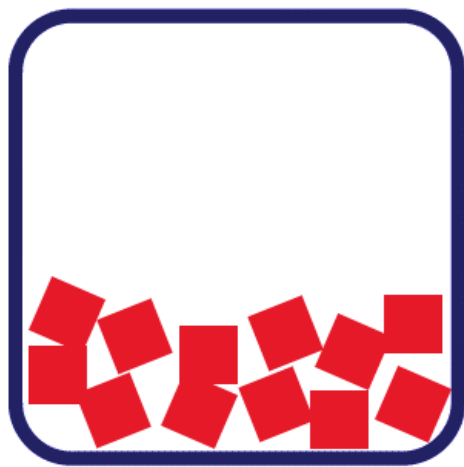


Figure 1.1.2 Particle arrangement inside a liquid

Gaseous Phase: the particles in a gas are separated and lack structure (Figure 1.1.3). Gas takes the shape of a container as gas has an indefinite shape. Gas has an indefinite volume—for example, water vapour.

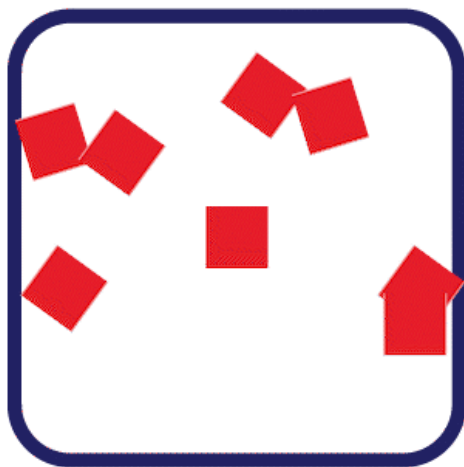


Figure 1.1.3 Particle arrangement inside a gas.

Chemical properties define how matter behaves in the presence of other matter in terms of its ability to undergo chemical transformations and change its form through chemical reactions. Understanding these properties is crucial for predicting and explaining the behaviour of matter in various chemical contexts. Chemical properties can not be observed or measured without changing the chemical identity of the substance. Take, for instance, the rusting of iron objects after exposure to moist air for a prolonged period (Figure 1.1.4). Here, the matter is iron, which has changed to iron oxide (rust). This means the initial chemical composition has changed while observing the chemical property, which is called corrosion¹. So, we say iron has the chemical property of corrosion. Another example is when wood burns, it changes to ashes. We say wood has the chemical property of flammability².

1. The ability of a material to deteriorate or react with its environment, often due to oxidation. The rusting of iron is an example of corrosion.
2. The tendency of a substance to undergo burning in the presence of oxygen



Figure 1.1.4 Example of corrosion observed in everyday life.
Image attribution: Old rust-covered motorcycle in an
abandoned building. AI-generated. © Wirestock –
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Changes in the Matter

If matter always stayed the same, chemistry would be rather boring. Fortunately, a major part of chemistry involves change. A physical change occurs when a sample of matter changes one or more of its physical properties. For example, a solid may melt (Figure 1.1.7), or alcohol in a thermometer may change volume as the temperature changes. A physical change does not affect the chemical composition of matter. A chemical change is the process of demonstrating a chemical property, such as the burning match in Figure 1.1.5 “Chemical Properties”. As the matter in the match burns, its chemical composition changes and new forms of matter with new physical properties are created. Note that chemical changes are frequently accompanied by physical changes, as the new matter will likely have different physical properties from the original matter.

Physical Change: Phase Change of Matter

A phase change refers to the transformation of a substance from one state of matter to another. The three primary states of matter are solid, liquid, and gas. Variations in temperature or pressure typically cause phase changes between these states.

Water is an excellent example to illustrate phase changes because it can exist in all three states.

1. **Melting (Solid to Liquid):** When you add heat to ice (solid water) at its melting point (0 degrees Celsius or 32 degrees Fahrenheit at standard atmospheric pressure), it absorbs energy and undergoes a phase change into liquid water.
2. **Freezing (Liquid to Solid):** The opposite of melting; freezing occurs when you remove heat from liquid water. The water molecules lose energy and arrange themselves into a more ordered, solid structure (ice).
3. **Vaporisation (Liquid to Gas):** Vaporisation is the process by which a liquid turns into a gas. It can occur in two forms: boiling and evaporation. Boiling happens at the liquid's boiling point. For water, boiling occurs at 100 degrees Celsius or 212 degrees Fahrenheit at standard atmospheric pressure, while evaporation can occur at any temperature.
4. **Condensation (Gas to Liquid):** Condensation is the reverse of vaporisation. It happens when a gas loses heat and transforms into a liquid. This occurs, for example, when water vapour in the air cools and forms dew on a cold surface.
5. **Sublimation (Solid to Gas):** Sublimation is the process by which a substance transitions directly from a solid to a gas without passing through the liquid phase. An example of this is dry ice (solid carbon dioxide) turning into carbon dioxide gas.
6. **Deposition (Gas to Solid):** Deposition is the reverse of sublimation. It occurs when a gas transforms directly into a solid without going through the liquid phase. Frost forming on

a cold surface is an example of deposition.

These phase changes are governed by the principles of thermodynamics, which describe the relationship between heat, energy, and matter (you will learn the basics of thermodynamics in chapter 7). The energy involved in these processes is typically associated with breaking or forming intermolecular forces between the particles of the substance.

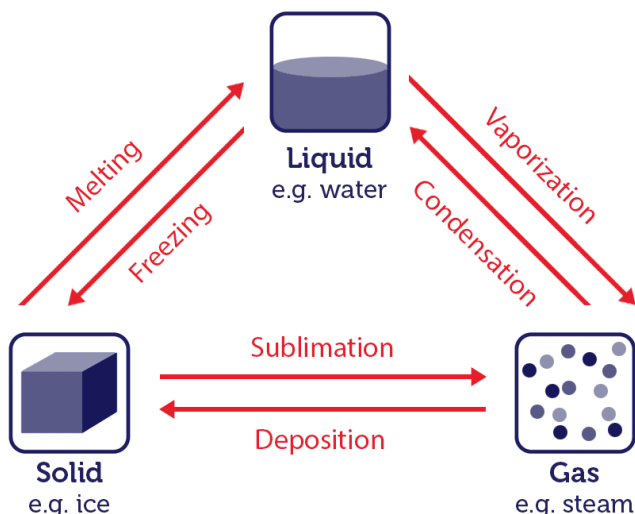


Figure 1.1.5 Water phase change.

Examples 1.1.1

Problems

Describe each process as a physical change or a chemical change.

1. Water in the air turns into snow.
2. A person's hair is cut.
3. Bread dough becomes fresh bread in an oven.

Solutions

1. Because the water is going from a gas phase to a solid phase, this is a physical change.
2. Your long hair is being shortened. This is a physical change.
3. Because of the oven's temperature, chemical changes are occurring in the bread dough to make fresh bread. These are chemical changes. (In fact, a lot of cooking involves chemical changes.)

Test Yourself

Identify each process as a physical change or a chemical change.

1. A fire is raging in a fireplace.
2. Water is warmed to make a cup of coffee.

Answers

1. Chemical change.
2. Physical change.

Classification of Matter

A sample of matter that has the same physical and chemical properties throughout is called a **substance**. Sometimes, the phrase *pure substance* is used, but the word *pure* isn't needed. This definition of substance is an example of how chemistry has a specific definition for a word that has been used in everyday language with a different, more vague definition. Here, we will use the term substance with its strict chemical definition.

Chemistry recognises two different types of substances: **elements** and **compounds**. An element is the simplest type of

chemical substance; it cannot be broken down into simpler chemical substances by ordinary chemical means. There are about 115 elements known to science, of which 80 are stable. The other elements are radioactive. Each element has its own unique set of physical and chemical properties. Examples of elements include iron, carbon, and gold.

A compound is a combination of more than one element. The physical and chemical properties of a compound are different from the physical and chemical properties of its constituent elements; that is, compounds behave as a substance that is completely different from the individual elements they are made from. There are over 50 million compounds known, and more are being discovered daily. Examples of compounds include water, penicillin, and sodium chloride (the chemical name for common table salt).

Elements and compounds are not the only ways matter can exist. We frequently encounter objects that are physical combinations of more than one element or compound. Physical combinations of more than one substance are called **mixtures**. There are two types of mixtures. A **heterogeneous mixture** is a mixture composed of two or more substances. It is easy to tell, sometimes by the naked eye, that more than one substance is present. A **homogeneous mixture** is a combination of two or more substances that is so intimately mixed that the mixture behaves as a single substance. Another word for a homogeneous mixture is a solution. Thus, a combination of salt and steel wool is a heterogeneous mixture because it is easy to see which particles of the matter are salt crystals and which are steel wool. On the other hand, if you take salt crystals and dissolve them in water, it is very difficult to tell that you have more than one substance present just by looking—even if you use a powerful microscope. The salt dissolved in water is a homogeneous mixture or a solution. Figure 1.1.6 displays examples of homogeneous and heterogenous mixtures.

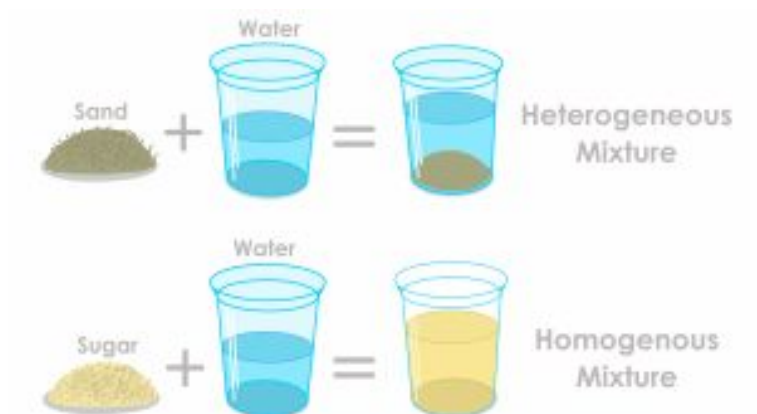


Figure 1.1.6 Different types of mixtures. Image attribution: Homogeneous, heterogeneous mixtures. Salt or sugar solution, sea. Sand depression with water in the glass. Solute, solvent molecules. Solid, liquid mix. Chemistry with explanations, Illustration Vector © LuckySoul – stock.adobe.com

There are other descriptors that we can use to describe matter, especially elements. We can usually divide elements into metals and nonmetals, and each set shares certain (but not always all) properties. A **metal** is an element that is solid at room temperature (although mercury is a well-known exception), is shiny and silvery, conducts electricity and heat well, can be pounded into thin sheets (a property called *malleability*), and can be drawn into thin wires (a property called *ductility*). A **nonmetal** is an element that is brittle when solid, does not conduct electricity or heat very well, and cannot be made into thin sheets or wires. Nonmetals also exist in a variety of phases and colours at room temperature. Some elements have properties of both metals and nonmetals and are called **metalloids**. We will see later how these descriptions can be easily assigned to various elements.

Problems

Identify the following combinations as heterogeneous mixtures or homogenous mixtures.

1. Soda water (carbon dioxide is dissolved in water).
2. A mixture of iron metal filings and sulphur powder (both iron and sulphur are elements).

Solutions

1. Because carbon dioxide is dissolved in water, we can infer from the behaviour of salt crystals dissolved in water that carbon dioxide dissolved in water is (also) a homogeneous mixture.
2. Assuming that the iron and sulphur are simply mixed together, it should be easy to see what is iron and what is sulphur, so this is a heterogeneous mixture.

Test Yourself

Are the following combinations homogeneous mixtures or heterogeneous mixtures?

1. The human body
2. An amalgam, a combination of some other metals dissolved in a small amount of mercury

Answers

1. Heterogeneous mixture.
2. Homogeneous mixture.

Watch the following simulation about the state of matter and phase changes.



One or more interactive elements has been excluded from this version of the text. You can view them online

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Key Takeaways

- Matter is anything that has mass and takes up space.
- Matter can be described in terms of physical properties and chemical properties.
- Physical properties and chemical properties of matter can change.
- Matter is composed of elements and compounds.
- Combinations of different substances are called mixtures.
- Elements can be described as metals, nonmetals, and semimetals.

Exercises

Practice Questions



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Transcript

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- Homogeneous, heterogeneous mixtures. Salt or sugar solution, sea. Sand depression with water in glass. Solute, solvent molecules. Solid, liquid mix. Chemistry with explanations, Illustration Vector © LuckySoul - stock.adobe.com

1.2 Expressing Numbers

Learning Objectives

- Express numbers using scientific notation.
- Apply the concept of significant figures to limit a measurement to the proper number of digits.
- Recognise the number of significant figures in a given quantity.
- Limit mathematical results to the proper number of significant figures.

Quantities have two parts: the number and the unit. The number tells “how many,” and the unit tells “the scale by which the number is measured”. It is important to be able to express numbers properly so that the quantities can be communicated accurately.

Scientific Notation

Standard notation is the straightforward expression of a number. Numbers such as **17**, **101.5**, and **0.00446** are expressed in standard notation. For relatively small numbers, standard notation is fine. However, for very large numbers, such as **306,000,000**, or for very small numbers, such as **0.000000419**, standard

notation can be cumbersome because of the number of zeros needed to place nonzero numbers in the proper position.

Scientific notation is an expression of a number using powers of **10**. Powers of **10** are used to express numbers that have many zeros (see Table 1.2.1).

Table 1.2.1 Powers of 10	
10^0	= 1
10^1	= 10
10^2	= 100 = 10×10
10^3	= 1,000 = $10 \times 10 \times 10$
10^4	= 10,000 = $10 \times 10 \times 10 \times 10$

and so forth. The raised number to the right of the **10** is the exponent, and it indicates the number of factors of 10 in the original number. Therefore, scientific notation is sometimes called *exponential notation*. The exponent’s value is equal to the number of zeros in the number expressed in standard notation.

Small numbers can also be expressed in scientific notation but with negative exponents; see Table 1.2.2.

Table 1.2.2 Powers of Negative 10.	
10^{-1}	= 0.1 = $\frac{1}{10}$
10^{-2}	= 0.01 = $\frac{1}{100}$
10^{-3}	= 0.001 = $\frac{1}{1,000}$
10^{-4}	= 0.0001 = $\frac{1}{10,000}$

and so forth. Again, the value of the exponent is equal to the number of zeros in the denominator of the associated fraction. A negative exponent implies a decimal number less than one.

A number is expressed in scientific notation by writing the first nonzero digit, then a decimal point, and then the rest of the digits. The part of a number in scientific notation that is multiplied by a power of **10** is called the coefficient. Then, determine the power of **10** needed to make that number into the original number and multiply the written number by the proper power of **10**.

For example, to write **79,345** in scientific notation. For example, **79,345** is written in scientific notation in this way:

$$79,345 = 7.9345 \times 10,000 = 7.9345 \times 10^4$$

Thus, the number in scientific notation is 7.9345×10^4 . For small numbers, the same process is used, but the exponent for the power of 10 is negative. For example:

$$0.000411 = 4.11 \times (1/10000) = 4.11 \times 10^{-4}$$

Typically, the extra zero digits at the end or the beginning of a number are not included.

Another way to determine the power of **10** in scientific notation is to count the number of places you need to move the decimal point to get a numerical value between **1** and **10**. The number of places equals the power of 10. This number is positive if you move the decimal point to the right and negative if you move the decimal point to the left.

Many quantities in chemistry are expressed in scientific notation. An important skill is learning how to correctly enter a number in scientific notation into your calculator. This will not be covered in this book because different models of calculators require different actions for properly entering scientific notation.

Examples 1.2.1

Problems

Express these numbers in scientific notation.

1. 306, 000
2. 0.00884
3. 2, 760, 000
4. 0.000000559

Solutions

1. The number 306, 000 is 3.06 times 100, 000, or 3.06 times 10^5 . In scientific notation, the number is 3.06×10^5 .
2. The number 0.00884 is 8.84 times $1/1000$, which is 8.84 times 10^{-3} . In scientific notation, the number is 8.84×10^{-3} .
3. The number 2, 760, 000 is 2.76 times 1, 000, 000, which is the same as 2.76 times 10^6 . In scientific notation, the number is written as 2.76×10^6 . Note that we omit the zeros at the end of the original number.
4. The number 0.000000559 is 5.59 times $(1/1000000)$, which is 5.59 times 10^{-7} . In

$$5.59 \times 10^{-7}$$

1

Express these numbers in scientific notation.

2. 0.0009706

Answers

1. 2.307×10^4

2. 9.706×10^{-4}

Significant Figures

If you use a calculator to evaluate the expression $(337 \div 217)$ **337217">**

you will get the following:

$$337 \div 217 = 1.55299539171$$

and so on for many more digits. Although this answer is correct, it is somewhat presumptuous. You start with two values that each have three digits, and the answer has twelve digits. That does not make much sense from a strict numerical point of view. Do we need to report all the digits that come after the decimal?

This concept of reporting the proper number of digits in a measurement or a calculation is called significant figures. Significant figures (sometimes called significant digits) represent the limits of what values of a measurement or a calculation we are sure of. The convention for a measurement is that the quantity reported should be all known values and the first estimated value. The conventions for calculations are discussed as follows.

In many cases, you will be given a measurement. How can you tell by looking what digits are significant? For example, the reported population of the United States is **306,000,000**. Does that

mean that it is exactly three hundred six million or is some estimation occurring?

The following conventions dictate which numbers in a reported measurement are significant and which are not significant:

1. Any nonzero digit is significant.
2. Any zeros between nonzero digits (i.e., embedded zeros) are significant.
3. Zeros at the end of a number without a decimal point (i.e., trailing zeros) are not significant; they serve only to put the significant digits in the correct positions. However, zeros at the end of any number with a decimal point are significant.
4. Zeros at the beginning of a decimal number (i.e., leading zeros) are not significant; again, they serve only to put the significant digits in the correct positions.

So, by these rules, the population figure of the United States has only three significant figures: the 3, the 6, and the zero between them. The remaining six zeros simply put the **306** in the millions position.

Examples 1.2.2

Problems

Give the number of significant figures in each measurement.

1. $36.7m$
2. $0.006606s$
3. $2,002kg$
4. $306,490,000$ people

Solutions

1. By rule 1, all nonzero digits are significant, so this measurement has three significant figures.
2. By rule 4, the first three zeros are not significant, but by rule 2 the zero between the sixes is; therefore, this number has four significant figures.
3. By rule 2, the two zeros between the twos are significant, so this measurement has four significant figures.
4. The four trailing zeros in the number are not significant, but the other five numbers are, so this number has five significant figures.

Test Yourself

Give the number of significant figures in each measurement.

1. $0.000601m$
2. $65.080kg$

Answers

1. three significant figures
2. five significant figures

How are significant figures handled in calculations? It depends on what type of calculation is being performed:

- If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column so that all numbers have significant figures in common. For example, if you were to add **1.2** and **4.41**, we note that the first number stops its significant figures in the tenth column, while the second number stops its significant figures in the hundredths column. We, therefore, limit our answer to the tenth column. We drop the last digit – the 1 – because it is not significant to the final answer.
- The dropping of positions in sums and differences brings up the topic of rounding. Although there are several conventions, in this text, we will adopt the following rule: the final answer should be rounded up if the first dropped digit is 5 or greater and rounded down if the first dropped digit is less than 5.
- If the operations being performed are multiplication or division, the rule is as follows: limit the answer to the number of significant figures that the data value with the least number of significant figures has. So, if we are dividing **23** by **448**, which have two and three significant figures each, we should limit the final reported answer to two significant figures (the lesser of two and three significant figures). The same rounding rules apply in multiplication and division as they do in addition and subtraction.

Problems

Express the final answer to the proper number of significant figures.

1. $101.2 + 18.702 = ?$
2. $202.88 - 1.013 = ?$
3. $76.4 \times 180.4 = ?$
4. $934.9 \div 0.00455 = ?$

Solutions

1. If we use a calculator to add these two numbers, we would get **119.902**. However, most calculators do not understand significant figures, and we need to limit the final answer to the tenth's place. Thus, we drop the **02** and report a final answer of **119.9** (rounding down).
2. A calculator would answer **201.867**. However, we have to limit our final answer to the hundredths place. Because the first number being dropped is **7**,

which is greater than **6**, we round up and report a final answer of **201.87**.

3. The first number has three significant figures, while the second number has four significant figures.

Therefore, we limit our final answer to three significant figures:

$$76.4 \times 180.4 = 13,782.56 = 13,800.$$

4. The first number has four significant figures, while the second number has three significant figures.

Therefore we limit our final answer to three significant figures:

$$934.9 \div 0.00455 = 205,472.5275 = 205,000.$$

Test Yourself

Express the final answer to the proper number of significant figures.

1. $3.445 + 90.83 - 72.4 = ?$
2. $22.4 \times 8.314 = ?$
3. $1.381 \div 6.02 = ?$

Answer

1. **21.9**
2. **186**
3. **0.229**

As you have probably realised by now, the biggest issue in

determining the number of significant figures in a value is zero. Is the zero significant or not? One way to unambiguously determine whether a zero is significant or not is to write a number in scientific notation. Scientific notation will include zeros in the coefficient of the number only if they are significant. Thus, the number 8.666×10^6 has four significant figures. However, the number 8.6660×10^6 has five significant figures. That last zero is significant; if it were not, it would not be written in the coefficient. So, when in doubt about expressing the number of significant figures in quantity, use scientific notation and include the number of zeros that are truly significant.

Key Takeaways

- Standard notation expresses a number normally.
- Scientific notation expresses a number as a coefficient times a power of 10.
- The power of 10 is positive for numbers greater than 1 and negative for numbers between 0 and 1.
- Significant figures in a quantity indicate the number of known values plus one place that is estimated.
- There are rules for which numbers in a quantity are significant and which are not significant.
- In calculations involving addition and subtraction, limit significant figures based on the rightmost place that all values have in common.
- In calculations involving multiplication and division, limit significant figures to the least number of significant figures in all the data values.

Practice Questions



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<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=607#h5p-47>

Transcript

1.3 Units and Conversions

Learning Objectives

- Learn the units that go with various quantities.
- Express units using their abbreviations.
- Make new units by combining numerical prefixes with units.
- Convert from one unit to another unit of the same type.

A number indicates “how much,” but the unit indicates “of what.” The “of what” is important when communicating a quantity. For example, if you were to ask a friend how close you are to Lake Erie and your friend says “six,” then your friend isn’t giving you complete information. Six *what*? Six miles? Six inches? Six city blocks? The actual distance to the lake depends on what units you use.

Chemistry, like most sciences, uses the International System of Units, or SI for short. The letters SI stand for the French “le Système International d’unités”. SI specifies certain units for various types of quantities based on seven fundamental units for various quantities (see Table 1.3.1).

Table 1.3.1 SI base units

Parameter	SI unit	Symbol
Time	seconds	s
Length	meter	m
Mass	kilograms	kg
Electric current	Ampere	A
Temperature	Kelvin	K
Amount	mole	mol
Luminous intensity	candela	cd

To express a quantity, you need to combine a number with a unit. If you have a length that is 2.4m , then you express that length as simply 2.4m . A time of $15,000\text{s}$ can be expressed as $1.5 \times 10^4 \text{ s}$ in scientific notation.

Sometimes, a given unit is not an appropriate size to easily express a quantity. For example, the width of a human hair is very small, and it doesn't make much sense to express it in metres. SI also defines a series of numerical prefixes that refer to multiples or fractions of a fundamental unit to make a unit more conveniently sized for a specific quantity. Table 1.3.2 "Multiplicative Prefixes for SI Units" lists the prefixes, their abbreviations, and their multiplicative factors. Some of the prefixes, such as kilo-, mega-, and giga-, represent more than one of the fundamental units, while other prefixes, such as centi-, milli-, and micro-, represent fractions of the original unit. Note, too, that once again, we are using powers of 10. Each prefix is a multiple of or fraction of a power of 10.

Table 1.3.2 Multiplicative prefixes for SI units.

Multiply	Prefix	Symbol
10	deca	d
10^2	hecto	h
10^3	kilo	k
10^6	Mega	M
10^9	Giga	G
10^{12}	Tera	T
10^{-1}	deci	d
10^{-2}	centi	c
10^{-3}	milli	m
10^{-6}	micro	μ
10^{-9}	nano	n
10^{-12}	pico	p

To use fractions to generate new units, simply combine the prefix with the unit itself; the abbreviation for the new unit is the combination of the abbreviation for the prefix and the abbreviation of the unit. For example:

- The kilometre (km) is $1,000 \times \textit{metre}$, or $1,000m$.
Thus, 5 kilometres ($5km$) is equal to $5,000m$.
- Similarly, a millisecond (ms) is $1/1000 \times \textit{second}$ one-thousandth of a second. Thus, $25ms$ is 25 thousandths of a second. You will need to become proficient in combining prefixes and units.
- You may recognise that one of our fundamental units, the kilogram, automatically has a prefix-unit combination, the kilogram. The word *kilogram* means $1,000g$.

In addition to the fundamental units, SI also allows for derived units based on a fundamental unit or units. There are many derived units

used in science. For example, the derived unit for area comes from the idea that area is defined as width times height. Because both width and height are lengths, they both have the fundamental unit of metre, so the unit of area is *metre* \times *metre* or, *metre*² (*m*²). This is sometimes referred to as “square metres.” A unit with a prefix can also be used to derive a unit for area, so we can also have *cm*², *mm*², or *km*² as acceptable units for area.

Volume is defined as length times width times height, so it has units of metre \times metre \times metre or metre³ (*m*³), sometimes referred to as “cubic metres.” The cubic metre is a rather large unit. So, another unit is defined that is somewhat more manageable: the litre (*L*). A litre is one-thousandth of a cubic metre. Prefixes can also be used with the litre unit so that we can speak of millilitres (one-thousandth of a litre; *mL*) and kilolitres (*1,000L*; *kL*).

Units not only are multiplied together but also can be divided. For example, if you are travelling at one metre for every second of time elapsed, your velocity is 1 metre per second, or *1m/s*. The word *per* implies division, so velocity is determined by dividing a distance quantity by a time quantity. Other units for velocity include kilometres per hour (*km/h*) or even micrometres per nanosecond (*μm/ns*). Later, we will see other derived units that can be expressed as fractions.

Examples 1.3.1

Problems

1. A human hair has a diameter of about $6.0 \times 10^{-5}m$. Suggest an appropriate unit for this measurement and write the diameter of a human hair in terms of that unit.
2. What is the velocity of a car if it goes $25m$ in $5.0s$?

Solutions

1. The scientific notation 10^{-5} is close to 10^{-6} , which defines the micro- prefix. Let us use micrometres as the unit for hair diameter. The number 6.0×10^{-5} can be written as 60×10^{-6} , and a micrometre is 10^{-6} m, so the diameter of a human hair is about $60\mu m$.
2. If velocity is defined as a distance quantity divided by a time quantity, then velocity is 25 metres/5.0 seconds. Dividing the numbers gives us $25 \div 5.0 = 5.0$, and dividing the units gives us metres/second or m/s. The velocity is $5.0m/s$.

Test Yourself

1. Express the volume of an Olympic-sized swimming pool, $2,500,000L$, in more appropriate units.
2. A common garden snail moves about 6.1 m in 30 min . What is its velocity in metres per minute (m/min)?

Answers

1. $2.5ML$
2. $0.203m/min$

Unit Conversions

In this section, we learn how to replace initial units with other units of the same type to get a numerical value that is easier to comprehend. We can use conversion factors to convert one unit to another.

- For instance, the conversion of milligrams to grams. From prefixes, we know $1000mg = 1g$

Both $1000mg$ and $1g$ refer to the same amount but in different units. Based on this relationship, we can create a pair of conversion factors:

$$\frac{1000mg}{1g} \quad \text{and} \quad \frac{1g}{1000mg}$$

- Express $150.0mg$ in grams?

We can calculate $150mg$ in grams using the conversion factor

$$\frac{1g}{1000mg} \text{ as follows:}$$

$$\frac{1g}{1000mg} \times 150.0mg = 0.1500g$$

- Express $0.275g$ in milligrams?

Here we can use the other conversion factor $\frac{1000mg}{1g}$ as follows:

$$\frac{1000mg}{1g} \times 0.275g = 275mg$$

When you observe the last two calculations, you can identify that the conversion factor used from the pair depends on the unit required to be converted.

Examples 1.3.2

Problems

1. Write conversion factors for the following units
 - a) Litres and millilitres
 - b) Ounces and grams
2. Convert 3.500lb pounds to grams ($1lb = 454g$).

Solutions

1. a) The relationship between two units can be expressed as $1L = 1000mL$. Therefore, the conversion factors will be:

$$\frac{1L}{1000mL} \text{ and } \frac{1000mL}{1L}.$$

b) The relationship between two units can be expressed as $1oz = 28.3g$. Therefore, the conversion factors will be:

$$\frac{1oz}{28.3g} \text{ and } \frac{28.3g}{1oz}.$$

2. Relationship between two units - $1lb = 454g$.

Conversion factors - $\frac{1lb}{454g}$ and $\frac{454g}{1lb}$. Only one of

these conversion factors is required to convert pounds to

grams. The conversion factor, $\frac{454g}{1lb}$, is required as it allows for the cancellation of the pound units, leaving the gram as the new unit.

$$\frac{454g}{1lb} \times 3.500lb = 1589g$$

Test Yourself

1. Write the conversion factors for centimetres to metres
2. Convert $250.0mg$ to grams

Answers

1. $1m = 100cm$. Therefore, $\frac{1m}{100cm}$ and

$$\frac{100cm}{1m}.$$

2. $\frac{1g}{1000mg} \times 250.0mg = 0.2500g$

Key Takeaways

- Numbers tell “how much,” and units tell “of what.”
- Chemistry uses a set of fundamental units and derived units from SI units.
- Chemistry uses a set of prefixes that represent multiples or fractions of units.
- Units can be multiplied and divided to generate new units for quantities.
- Units can be converted to other units using the proper conversion factors.
- Conversion factors are constructed from equalities that relate two different units.

Exercises

Practice Questions



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Transcript

1.4 Temperature and Density

Learning Objectives

- Learn about the various temperature scales that are commonly used in chemistry.
- Define density and use it as a conversion factor.

Temperature

One of the fundamental quantities in science is temperature. Temperature is a measure of the average amount of energy of motion, or *kinetic energy*¹, a system contains. Temperatures are expressed using scales that use units called degrees, and there are several temperature scales in use. One of the commonly used temperature scales is the *Fahrenheit scale* (symbolised by °F and spoken as “degrees Fahrenheit”). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32°F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212°F.

Science also uses other scales to express temperature. The

1. Kinetic energy is a form of energy associated with the motion of an object

Celsius scale (symbolised by °C and spoken as “degrees Celsius”) is a temperature scale where 0°C is the freezing point of water and 100°C is the boiling point of water; the scale is divided into 100 divisions between these two landmarks and extended higher and lower. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined as follows:

$$^{\circ}\text{C} = (^{\circ}\text{F} - 32) \times \frac{5}{9}$$

$$^{\circ}\text{F} = \left(^{\circ}\text{C} \times \frac{9}{5}\right) + 32$$

Using these formulas, we can convert from one temperature scale to another. The number 32 in the formulas is exact and does not count in significant figure determination.

Example 1.4.1

Problems

1. What is 98.6°F in degrees Celsius?
2. What is 25.0°C in degrees Fahrenheit?

Solutions

1. Using the first formula from above, we have:

$$^{\circ}\text{C} = (98.6 - 32) \times \frac{5}{9} = 66.6 \times \frac{5}{9} = 37.0^{\circ}\text{C}$$

2. Using the second formula from above, we have:

$$^{\circ}\text{F} = \left(25.0 \times \frac{9}{5}\right) + 32 = 45.0 + 32 = 77.0^{\circ}\text{F}$$

Test Yourself

1. Convert 0°F to degrees Celsius.
2. Convert 212°C to degrees Fahrenheit.

Answers

1. -17.8°C
2. 414°F

The fundamental unit of temperature (another fundamental unit of science, bringing us to four) in SI is the kelvin (K). The Kelvin temperature scale (note that the name of the scale capitalises the word *Kelvin*, but the unit itself is lowercase) uses degrees that are the same size as the Celsius degree, but the numerical scale is shifted up by 273.15 units. That is, the conversion between the Kelvin and Celsius scales is as follows:

$$\begin{aligned} \text{K} &= ^\circ\text{C} + 273.15 \\ ^\circ\text{C} &= \text{K} - 273.15 \end{aligned}$$

For most purposes, it is acceptable to use 273 instead of 273.15. Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as “two hundred ninety-five kelvins” and not “two hundred ninety-five degrees Kelvin.”

The reason that the Kelvin scale is defined this way is because there exists a minimum possible temperature called absolute zero. The Kelvin temperature scale is set so that 0 K is absolute zero, and the temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example.

Example 1.4.2

Problem

If normal room temperature is 72.0°F, what is room temperature in degrees Celsius and kelvins?

Solution

First, we use the following formula to determine the temperature in degrees Celsius:

$$^{\circ}\text{C} = (72.0 - 32) \times \frac{5}{9} = 40.0 \times \frac{5}{9} = 22.2^{\circ}\text{C}$$

Then, we use the appropriate formula above to determine the temperature in the Kelvin scale:

$$\text{K} = 22.2^{\circ}\text{C} + 273.15 = 295.4 \text{ K}$$

So, the room temperature is about 295 K.

Test Yourself

What is 98.6°F on the Kelvin scale?

Answer

310.2 K

Figure 1.4.1 “Fahrenheit, Celsius, and Kelvin Temperatures” compares the three temperature scales. Absolute zero in each scale is -459°F , -273°C , and 0 K. Water freezes at 32°F , 0°C , and 273 K. 180°F equals 100°C , which equals 100 kelvins. Water boils at 212°F , 100°C , and 373 K. Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practising chemist expresses laboratory-measured temperatures with the Fahrenheit scale.

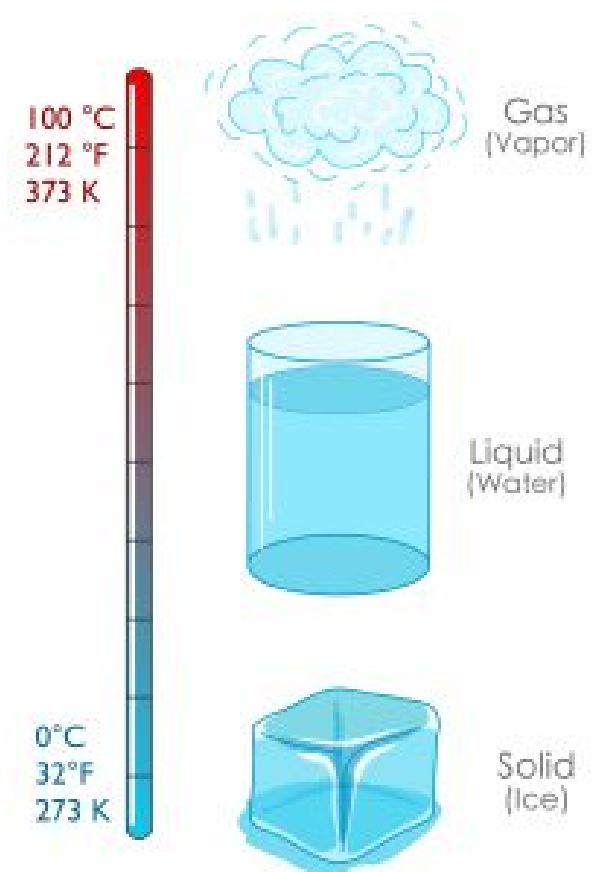


Figure 1.4.1 “Fahrenheit, Celsius, and Kelvin Temperatures.” A comparison of the three temperature scales with respect to phase changes of water. Image attribution: Water, solid, ice cube, liquid, gas, vapor. State of matter. Change of state. Chemistry, physics. Phase fluid. Freeze, melt, evaporation. Temperature Fahrenheit, kelvin, celsius. Vector illustration © LuckySoul-stock.adobe.com

Density

Density is a physical property that is defined as a substance’s mass divided by its volume, as shown by the following equation:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \text{ or } d = \frac{m}{V}$$

In simpler terms, density describes how much mass is contained in a given volume. Objects with the same volume but different masses have different densities. This is because there is more or less mass packed into the same volume.

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units: mass and volume. That means that density overall has derived units, just like velocity. Common units for density include g/mL, g/cm³, g/L, kg/L, and even kg/m³. Densities for some common substances are listed in Table 1.4.1, “Densities of Some Common Substances”.

Table 1.4.1 Densities of Some Common Substances.

Substance	Density (g/mL or g/cm ³)
Water	1.0
Gold	19.3
Mercury	13.6
Air	0.0012
Cork	0.22–0.26
Aluminum	2.7
Iron	7.87

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm³. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If

you multiply the given volume by the known density (from Table 1.4.1 “Densities of Some Common Substances”), the volume units will cancel and leave you with mass units, telling you the mass of the sample is:

$$7.88 \text{ cm}^3 \times 2.7 \text{ g/cm}^3 = 21 \text{ g of aluminum}$$

where we have limited our answer to two significant figures.

Example 1.4.3

Problem

What is the mass of 44.6 mL of mercury?

Solution

Use the density from Table 1.4 “Densities of Some Common Substances” as a conversion factor to go from volume to mass:

$$44.6 \text{ mL} \times 13.6 \text{ g/mL} = 607 \text{ g}$$

The mass of the mercury is 607 g.

Test Yourself

What is the mass of 25.0 cm^3 of iron?

Answer

197 g

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

$$d = 19.3 \frac{\text{g}}{\text{mL}} = \frac{19.3 \text{ g}}{1 \text{ mL}}$$

Although this was not previously pointed out, it can be assumed that there is a 1 in the denominator.

That is, the density value tells us that we have 19.3 grams for every 1 millilitre of volume, and the 1 is an exact number. When we want to use density to convert from mass to volume, the numerator and denominator of density need to be switched – in other words, we must take the *reciprocal* of the density. In so doing, we move not only the units but also the numbers, as can be seen below:

$$\frac{1}{d} = \frac{1 \text{ mL}}{19.3 \text{ g}}$$

This reciprocal density is still a useful conversion factor, but now the mass unit will cancel, and the volume unit will be introduced. Thus, if we want to know the volume of 45.9 g of gold, we would set up the conversion as follows:

$$45.9 \text{ g} \times \frac{1 \text{ mL}}{19.3 \cancel{\text{ g}}} = 2.38 \text{ mL}$$

Note how the mass units cancel, leaving the volume unit, which is what we're looking for.

Example 1.4.4

Problem

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/cm^3 , what is the volume of the cork?

Solution

To use density as a conversion factor, we need to take the reciprocal so that the mass unit of density is in the denominator. Taking the reciprocal, we find:

$$\frac{1}{d} = \frac{1 \text{ cm}^3}{0.22 \text{ g}}$$

We can use this expression as the conversion factor. So:

$$3.78 \cancel{\text{g}} \times \frac{1 \text{ cm}^3}{0.22 \cancel{\text{g}}} = 17.2 \text{ cm}^3$$

Test Yourself

What is the volume of 3.78 g of gold?

Answer

0.196 cm³

Care must be taken with density as a conversion factor. Make sure the mass units are the same, or the volume units are the same before using density to convert to a different unit. Often, the unit of the given quantity must be first converted to the appropriate unit before applying density as a conversion factor.

Key Takeaways

- Chemistry uses the Celsius and Kelvin scales to express temperatures.
- A temperature on the Kelvin scale is the Celsius temperature plus 273.15.
- The minimum possible temperature is absolute zero and is assigned 0 K on the Kelvin scale.
- Density relates a substance's mass and volume.

- Density can be used to calculate volume from a given mass or mass from a given volume.

Exercises

Practice Questions



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Transcript

Media Attributions

- Water, solid, ice cube, liquid, gas, vapor. State of matter. Change of state. Chemistry, physics. Phase fluid. Freeze, melt, evaporation. Temperature Fahrenheit, kelvin, celsius. Vector illustration © LuckySoul-stock.adobe.com

PART II

CHAPTER 2: ATOMS AND THE PERIODIC TABLE

2.1 Atoms: The Building Blocks of Everything

Learning Objectives

1. State the modern atomic theory.
2. Learn how atoms are constructed.
3. Define the terms atoms, molecules and compounds.

Atoms and Elements

Within the chemical world, **atoms** are the essential building blocks, composing the many **molecules** and compounds we interact with. All atoms consist of a small, positively charged nucleus that is made up of positively charged **protons**, as well as **neutrons**, which themselves have no charge. This is surrounded by a larger, negatively charged **electron** cloud (Figure 2.1.1). They are too small to be observed by the naked eye.

The concept that atoms play a fundamental role in chemistry is formalised by modern atomic theory, first stated by John Dalton, an English scientist, in 1808. It consists of three parts:

1. All matter is composed of atoms.
2. Atoms of the same element are the same; atoms of different elements are different.
3. Atoms combine in whole-number ratios to form compounds.

These concepts form the basis of chemistry.

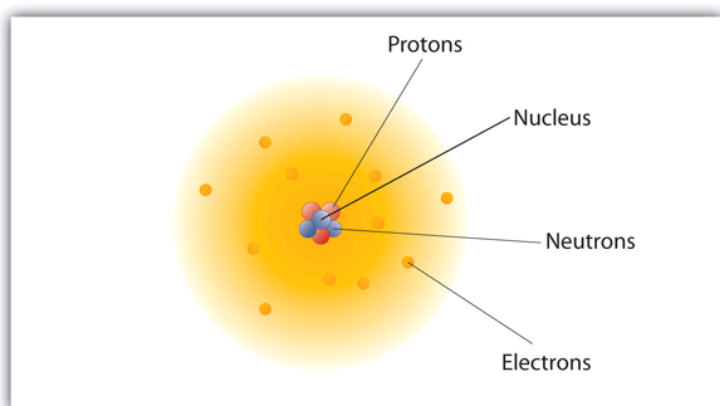


Figure 2.1.1: “The Structure of the Atom.” Atoms have protons and neutrons in the centre, making the nucleus, while the electrons orbit the nucleus in the surrounding electron cloud. Image attribution: “Nucleus” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike)

Although the word *atom* comes from a Greek word that means “indivisible,” we now understand that atoms themselves are composed of smaller parts called **subatomic particles**. The subatomic particles present in an atom’s nucleus—protons and neutrons—are collectively known as nucleons.. Protons and neutrons are almost identical in mass, whereas electrons are much lighter. Protons hold a positive (+1) charge, while electrons hold a negative (-1) charge. The properties of the subatomic particles are given in Table 2.1.1. Atoms are generally **neutral species** due to the presence of an equal number of protons and electrons.

Subatomic particle	Symbol	Charge
Proton	p	+1
Neutron	n	0
Electron	e	−1

Atoms come in many different “elements” — depending on the number of protons present. Each element has a unique name and a **chemical symbol**. We will learn how to read the periodic table in 2.2 ‘Electronic Configuration’, to determine the name of a given **element**. Examples of element symbols are given in Table 2.1.2.

Element
Hydrogen
Carbon
Calcium
Silver
Copper

Elements act as the foundations of all matter. Some elements naturally exist as multiatomic, for instance, S_8 , O_2 , and Cl_2 while others such as **He**, **Ar**, and, **Ne** naturally stand alone. This can be predicted by observing the electrons present, which will be explained in greater detail in the next chapter.

Atomic Number and Mass Number

As atoms can have differing numbers of **subatomic particles** we utilise atomic and mass numbers to determine their amount. The **atomic number** (Z) is the number of protons present in the nucleus of an atom.

Atomic number (Z) = Number of protons

The **mass number** (A) of an atom is the sum of the number of protons and neutrons present in the nucleus of an atom.

Mass number (A) = Number of protons + Number of neutrons

Let's observe how these formulae can be used to determine the atomic number and mass number of a standard oxygen atom, as illustrated in Figure 2.1.2. Within the oxygen atom, we have 8 protons, 8 neutrons and 8 electrons.

Atomic number (Z) = Number of protons = 8

Mass number (A) = Number of protons + Number of neutrons = 8 + 8 = 16.

It can, therefore, be seen that a standard oxygen atom has an atomic number of 8 and a mass number of 16. We represent this as $^{16}_8\text{O}$.

The relationship between atomic numbers, mass numbers and the number of protons and neutrons present within an atom allows for a variety of problem-solving equations. For instance, one can work out the atomic number of an atom from the number of neutrons and the mass number alone.

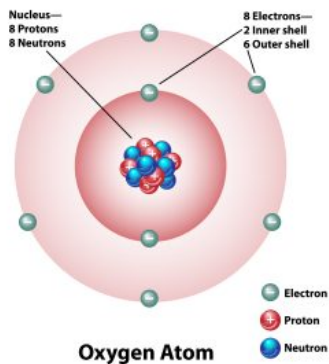


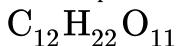
Figure 2.1.2: Oxygen atom diagram. Image attribution: Diagram of an oxygen atom with nucleus and inner and outer shells. Protons, neutrons, and electrons are labeled. © O Sweet Nature – stock.adobe.com

Molecules and Compounds

While this chapter focuses on atoms by themselves, important terminology should be established. The terms atoms, molecules and compounds can't be used interchangeably. A molecule is a group of two or more atoms combined chemically and functions as a unit. The simplest molecule that can exist is a **diatomic molecule** that contains two atoms, such as O_2 and H_2 . **Compounds** contain different types of atoms in fixed proportions. Compounds and molecules are uncharged, neutral species.

Chemical formulas consist of chemical symbols of the elements present in the compound and numerical subscripts, which demonstrate the number of atoms of each element involved in the formation of the compound.

- The chemical formula for the compound sucrose is

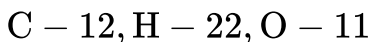


This formula shows that sucrose contains three different elements:

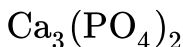


Total number of atoms: 45

Number of each type of atom:



- The chemical formula for the compound calcium phosphate is



This formula shows that one unit of calcium phosphate is composed of three different elements: $Ca\ P\ O$

Total number of atoms: 13

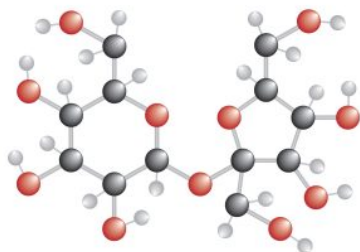


Figure 2.1.3: Ball-and-stick model of Sucrose compound. Image attribution: Molecular formula of sucrose. Sucrose or tea sugar is a disaccharide formed by the combination of a glucose and a fructose molecule. © Firat - stock.adobe.com

Number of each type of atom: $\text{Ca} - 3, \text{P} - 2, \text{O} - 8$
(Note: as there are two phosphate ions (PO_4), for O and P multiply the subscript number by the number written after the parenthesis.)

Key Takeaways

- Atoms are small particles comprised of positively charged protons, neutral neutrons, and negatively charged electrons.
- Atoms come in many different elements, denoted by symbols found on the periodic table (such as H or Cu).
- The atomic number of a particle reflects the number of protons it contains, while the mass number is the full amount of protons and neutrons added together.
- Chemical formulae tell you how many of which elements are involved in the makeup of a compound

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

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- Diagram of an oxygen atom with nucleus and inner and outer shells. Protons, neutrons, and electrons are labeled. © O Sweet Nature - stock.adobe.com is licensed under a All Rights Reserved license
- Molecular formula of sucrose. Sucrose or tea sugar is a disaccharide formed by the combination of a glucose and a fructose molecule. © Firat - stock.adobe.com is licensed under a All Rights Reserved license

2.2 Electronic Configuration

Learning Objectives

1. Learn how electrons are organised in atoms.
2. Represent the organisation of electrons by an electron/electronic configuration.
3. Understand the implications of paired and unpaired electrons within atoms.

Electron shells

Understanding how **electrons** arrange themselves within an atom is essential to predicting chemical bonding and interactions between atoms.

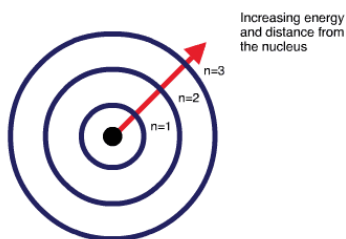


Figure 2.2.1: Increasing energy and distance from the nucleus.

Space, where electrons move around the nucleus inside an atom, is divided into subspaces known as shells, subshells and orbitals.

The primary energy levels are called shells. The energy of the electron shells increases with distance from the nucleus.

Therefore, the electrons closest to the nucleus have lower energy than the electrons located further

away. Electron shells are named outward from the nucleus as **1, 2, 3, 4** and so on.

Electron shells have a maximum capacity of electrons which can occupy the entire shell at one time. This rule is known as the “2n Squared Rule”, where the *Shell Capacity* = $2n^2$, with n equalling the shell number. As such, the first shell ($n = 1$) can hold 2 electrons, the second shell ($n = 2$) can hold 8 and so on.

While this rule is useful for understanding the capacity of a shell, it doesn’t predict the order in which electrons enter shells. Rather, electrons naturally occupy the lowest energy states possible. To understand that, we must analyse subshells and orbitals.

Electron subshell and orbitals

Within electron shells sit subshells. Subshells are labelled and ordered as *s*, *p*, *d* and *f*. Each subshell carries 2 electrons for each orbital present. An *s* subshell has 1 orbital and can carry 2 electrons, a *p* subshell has 3 orbitals and can carry 6 electrons, and so on. The number of subshells in a shell is equal to the shell number.

For instance, in the 3rd ($n=3$) electron shell, 3 subshells (*s*, *p*, *d*) are present. If we add the amount of electrons within each subshell, we get 18 according to the 2n Squared Rule.

Shell number (n)	Subshells present	Number of orbitals	Number of electrons
1	<i>s</i>	1	2
2	<i>s</i>	1	2
	<i>p</i>	3	6

3	<i>s</i>	1	2
	<i>p</i>	3	6
	<i>d</i>	5	10
4	<i>s</i>	1	2
	<i>p</i>	3	6
	<i>d</i>	5	10
	<i>f</i>	7	14

Table 2.2.1: Electronic Shells, Subshells and Orbitals.

Schrödinger’s Model: Making Sense of Uncertainty.

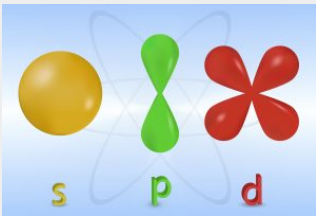


Figure 2.2.2: The shapes of atomic orbitals. Image attribution: Atomic orbitals © extender_01 – stock.adobe.com

Orbitals are a product of the Schrödinger wave-mechanical model of electrons. Electrons in this model are not seen as physical particles but as energy waves surrounding the nucleus of an atom. Quantum mechanics tells us that it

is impossible for us to know the location and velocity of where the electron is simultaneously. As such, the orbital is a prediction of where this energy can be likely found within an atom. For most of our investigation within this textbook, considering electrons as single particles simplifies explanations. Although we will bring up the orbital explanation if necessary.

Within an electron orbital, the two electrons enter spin states: one spin up, and the other spin down.

Orbitals have distinct shapes, for instance. *s* orbitals are spherical and *p* orbitals are dumbbell shaped. This has further implications for predicting the bonding and stability of compounds. As orbitals are probability distributions of electrons, these can be added and subtracted when atoms bond. For double bonds, this causes the formation of pi (π) orbitals. *This, however, is beyond the scope of what needs to be known in this chapter.*

Predicting electron configurations

Electron subshells are filled in order of increasing energy. Figure 2.2.3 showcases the Aufbau principle: the order in which subshells fill with electrons depending on their energy. Beginning from the top left $1s$ shell, the order proceeds

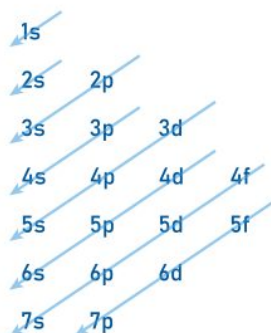


Figure 2.2.3: Aufbau Principle. Order to fill electron orbitals. Image attribution: Aufbau order of orbitals for feeding in electrons. Electron configuration. Aufbau principle. Filling orbitals in atom. Vector illustration isolated on white background. © SAMYA – stock.adobe.com

$2s$, $2p$, $3s$, $3p$, $4s$, $3d$, $4p$, $5s$ and so on. Note how the $4s$ subshell fills before the $3d$ subshell.

Electrons fill into orbitals of a subshell so that each orbital acquires one electron before any orbital acquires a second electron. All single electrons must have the same spin. Electron energy diagrams allow us to see this in action. Let us demonstrate the **electron configuration** of a standard boron atom, which has 5 electrons.

In Figure 2.2.4, we have drawn the first five electron subshells ($1s$ to $3p$) in order of ascending energy levels with their respective orbitals. Note how the s subshells have only 1 orbital present, whereas the p subshells hold 3 each.

Let's take our boron example. In a standard boron atom, we have 5 electrons to assign to orbitals. Beginning from the lowest energy subshells, two electrons (represented by up and down arrows) can be placed into the $1s$ subshell and two more into the $2s$ subshell. The leftover electron can be placed into one of the $2p$ orbitals.

We can now describe the electronic composition of boron, being $1s^2 2s^2 2p^1$: 2 electrons in $1s$, 2 in $2s$ and 1 in $2p$.

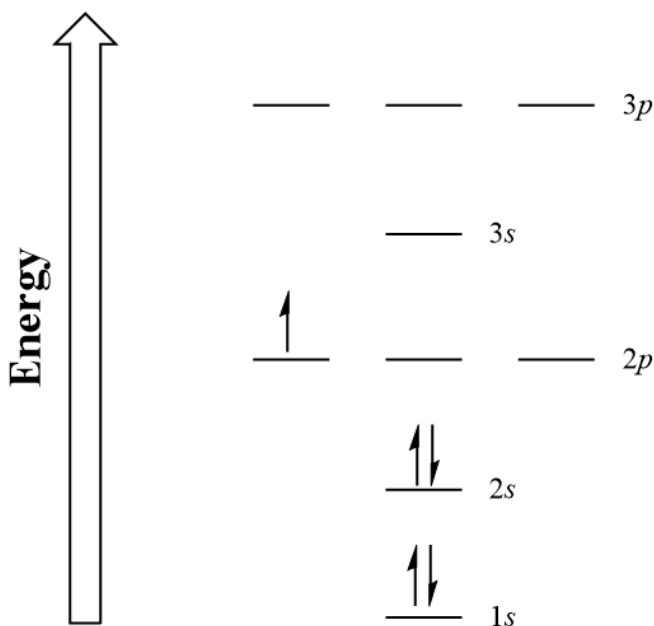


Figure 2.2.4: Boron Electronic Energy Diagram Image attribution: "Generic Energy Diagram of Orbitals in a Multi-Electron Atom" by David W. Ball and Jessie A. Key © CC BY (Attribution). Transcript.

Moving across the periodic table, we can see how an additional electron is filled into $2p$ orbitals for each subsequent element (see Figure 2.2.5). For oxygen, we can finally pair up electrons with their opposite spin and fill one of the $2p$ orbitals.

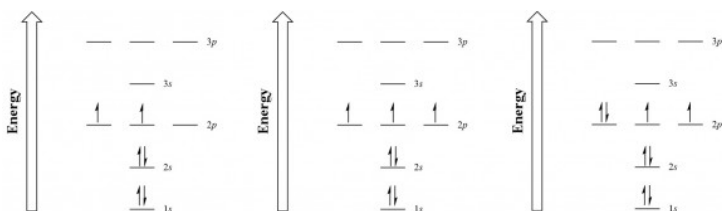


Figure 2.2.5: Carbon (left-side), Nitrogen (middle) and Oxygen (right-side) Electronic Energy Diagrams. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License, Transcript.

Electronic configurations¹ are vital to understanding atomic bonding and the structure of the periodic table. As we will discuss in chapter 3.1 ‘The Octet Rule and Lewis Dot Diagrams’, **valence electrons** heavily dictate how many bonds an atom can achieve. Atoms naturally generate stable bonds by filling their outermost shell full of electrons. In the case of oxygen, two more electrons are needed in the **2p** subshell to create a full 2nd shell – allowing electrons to form stable chemical bonds with two different atoms. For elements later down the periodic table, particularly the transition metals, **ions** that have extra or fewer electrons than normal may have the same amount of valence electrons due to the filling order. The periodic table, as we will see next chapter, can be divided up into subshell blocks, allowing us to predict chemical properties.

1. Electronic configuration and electron configuration are the same thing.

Table 2.2.2: The electron configuration of the first ten elements on the periodic table.

Element	Atomic number	Electron configuration
H	1	$1s^1$
He	2	$1s^2$
Li	3	$1s^2 2s^1$
Be	4	$1s^2 2s^2$
B	5	$1s^2 2s^2 2p^1$
C	6	$1s^2 2s^2 2p^2$
N	7	$1s^2 2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$
Ne	10	$1s^2 2s^2 2p^6$

Key Takeaways

- Electrons occupy orbitals found within subshells.
- Subshells have an increasing number of orbitals as the shell number increases.
- Electrons fill orbitals from the lowest energy upwards.
- The electronic configuration is predicted using the

Aufbau Principle, with the number of unpaired valence electrons vital to bonding.

- Unless ionic or excited, the electron configuration of an element will always be the same.

Exercises

Practice Questions



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Transcript

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- Aufbau order of orbitals for feeding in electrons. Electron configuration. Aufbau principle. Filling orbitals in atom. Vector illustration isolated on white background. © SAMYA - stock.adobe.com is licensed under a All Rights Reserved license

2.3 The Periodic Table

Learning Objectives

1. Relate the electron configurations of the elements to the shape of the periodic table.
2. Determine the expected electron configuration of an element by its place on the periodic table.

To organise the various elements in chemistry, the periodic table arranges elements in order of increasing atomic number. The rows and columns in the periodic table are known as periods and groups, respectively (see Figure 2.3.1).

Periodic Table of the Elements

The periodic table displays elements from Hydrogen (H) to Oganesson (Og). Each element cell includes its atomic number, symbol, name, and atomic weight. The table is organized into groups (columns) and periods (rows). Key features include:

- Groups:** 1 (IA), 2 (IIA), 3 (IIIB), 4 (IVB), 5 (VB), 6 (VIB), 7 (VIIB), 8 (VIII), 9 (VIII), 10 (VIII), 11 (IB), 12 (IIB), 13 (IIIA), 14 (IVB), 15 (VA), 16 (VIA), 17 (VIIA), 18 (VIIIA).
- Periods:** 1, 2, 3, 4, 5, 6, 7, 8, 9, 10, 11, 12, 13, 14, 15, 16, 17, 18.
- Element Categories:**
 - Alkali metals (Group 1)
 - Alkaline earth metals (Group 2)
 - Transition metals (Groups 3-10)
 - Post-transition metals (Groups 11-12)
 - Metalloids (Groups 13-14)
 - Nonmetals (Groups 15-16)
 - Noble gases (Group 18)
- Legend:**
 - State of matter (color): Solid (blue), Liquid (orange), Gas (green), Unknown (grey).
 - Element category (color): Alkali metals (red), Alkaline earth metals (orange), Transition metals (blue), Post-transition metals (green), Metalloids (yellow), Nonmetals (light green), Noble gases (pink).

Figure 2.3.1: Periodic Table of the Elements. Image attribution: Colorful Periodic Table of the Elements – shows atomic number, symbol, name, atomic weight, electrons per shell, state of matter and element category © Humdan – stock.adobe.com. View an accessible periodic table online.

Why does the periodic table have the structure it does? The answer is rather simple if you understand electron configurations: the shape of the periodic table mimics the filling of the subshells with electrons.

Dividing the table by electronic subshells

Let us start with **H** and **He**. Their electron configurations are $1s^1$ and $1s^2$, respectively; with **He**, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (see Figure 2.3.2 “The 1s Subshell”).

Figure 2.3.2: “The 1s Subshell.” H and He represent the filling of the 1s subshell. Image attribution: “The 1s Subshell” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

A diagram of a periodic table with the 2s orbital highlighted in green. An arrow points to the green box with the label $2s$.

Figure 2.3.3: “The 2s Subshell.” In Li and Be, the 2s subshell is being filled. Image attribution: “The 2s Subshell” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

A periodic table grid showing the first four periods. The 2p subshell is highlighted in yellow, corresponding to the second period elements from Boron (B) to Neon (Ne). An arrow labeled $2p$ points to the first yellow cell (Boron).

[illegible]

The **3s** subshell is then filled. The elements when this subshell is being filled, **Na** and **Mg**, are on the left side of the periodic table (Figure 2.3.5 “The 3s Subshell”).

A blank periodic table grid is shown. The 3s orbital is highlighted in yellow. An arrow points to the 3s orbital with the label "3s".

[illegible]

Figure 2.3.5: “The 3s Subshell.” Now the 3s subshell is being occupied. Image attribution: “The 3s Subshell” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

Next, the **3p** subshell is filled with the next six elements (Figure 2.3.6 “The 3p Subshell”).

[illegible]

Figure 2.3.6: “The 3p Subshell.” Next, the 3p subshell is filled with electrons. Image attribution: “The 3p Subshell” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

Recall 2.2 ‘Electronic Configuration’ and the electron subshell filling orders. Instead of filling the **3d** subshell next, electrons go into the **4s** subshell, which consists of **K** and **Ca** (Figure 2.3.7 “The 4s Subshell”).

Figure 2.3.7: “The 4s Subshell.” The 4s subshell is filled before the 3d subshell. This is reflected in the structure of the periodic table. Image attribution: “The 4s Subshell” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

After the **4s** subshell is filled, the **3d** subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table, which consists of **Sc** to **Zn** (Figure 2.3.8 “The 3d Subshell”).

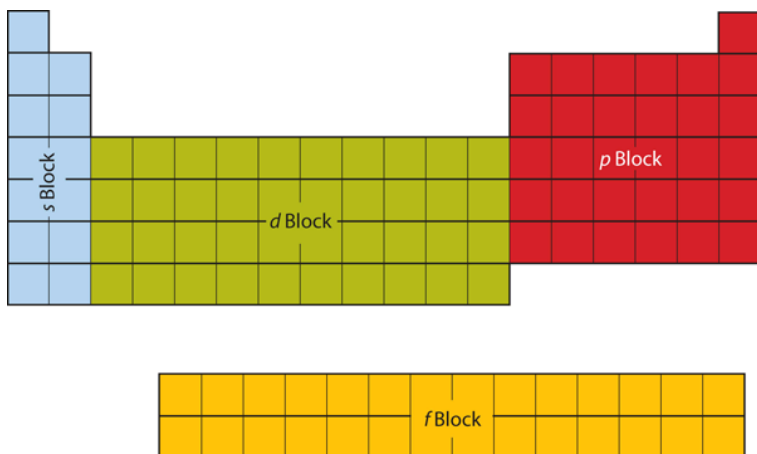


Figure 2.3.9: “Blocks on the Periodic Table.” The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section. Image attribution: “Blocks on the Periodic Table” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called *core electrons*.) The valence electrons largely control the expected chemistry and reactive properties of an atom. To illustrate, we find that in each column of the periodic table, the valence shell’s electron configuration is the same. Take the elements in the first column of the periodic table: **H**, **Li**, **Na**, **K**, **Rb**, and **Cs**. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

Table 2.3.1 Electron Configurations of Elements in the First Column of the Periodic Table

Element	Electron Configuration
H	$1s^1$
Li	$1s^2 2s^1$
Na	$[\text{Ne}]3s^1$
K	$[\text{Ar}]4s^1$
Rb	$[\text{Kr}]5s^1$
Cs	$[\text{Xe}]6s^1$

They all have a similar electron configuration in their valence shells: a single *s* electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar chemistry — *and they do*. The organisation of electrons in atoms explains not only the shape of the periodic table but also the fact that elements in the same column of the periodic table have similar chemistry.

Dividing the table by chemical properties

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the *s* and *p* blocks. In the *d* and *f* blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected. To this extent, the groups can be further classified as seen in Table 2.3.2.

Table 2.3.2 Periodic Table Common Groups and Properties

Group	Name	Examples	Chemical properties
1 (<i>IA</i>)	Alkali metals	Li, Na, K	Soft and shiny metals, highly reactive with water
2 (<i>IIA</i>)	Alkaline earth metals	Be, Mg, Ca	Soft and shiny metals, moderately reactive with water
17 (<i>VIIA</i>)	Halogens	F, Cl, Br	Generally, reactive elements, exist as gases at room temperature
18 (<i>VIIIA</i>)	Noble gases	He, Ne, Ar	Exist as gases, unreactive elements

Based on selected physical properties, elements are further classified into metals and nonmetals. Metals are located on the left side of the periodic table, and nonmetals are located on the right. Metals are good conductors of electricity and heat; exist as solids at room temperature (except mercury); are **ductile**; **malleable**; have shiny appearances (metallic lustre); and have high density and high melting points.

Elements known as metalloids exhibit both metallic and nonmetallic properties. Metalloids are located between metals and nonmetals in the periodic table.

Key Takeaways

- The arrangement of electrons in atoms is responsible for the shape of the periodic table.
- Electron configurations can be predicted by the position of an atom on the periodic table.
- Elements within the same group or column have similar valence electron shell configurations – often exhibiting similar properties as a result.

Exercises

Practice Questions



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Transcript

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2.4 Periodic Trends

Learning Objectives

1. Understand how certain properties of atoms vary based on their relative position on the periodic table.

Arranging the periodic table by **electron configuration** allows us to observe trends within elements across groups and periods. Let us discuss some important chemical properties which can be observed.

Atomic radii

The atomic radius is an indication of the size of an atom. While atoms don't have a size in the traditional sense (due to the fluctuating position of electrons), they behave as if they have a radius, particularly with reference to bonding. Atomic radii increase from the top to the bottom of a group of the periodic table. As n increases when proceeding down a group, orbitals become larger, increasing atomic radii (see Figure 2.4.1).

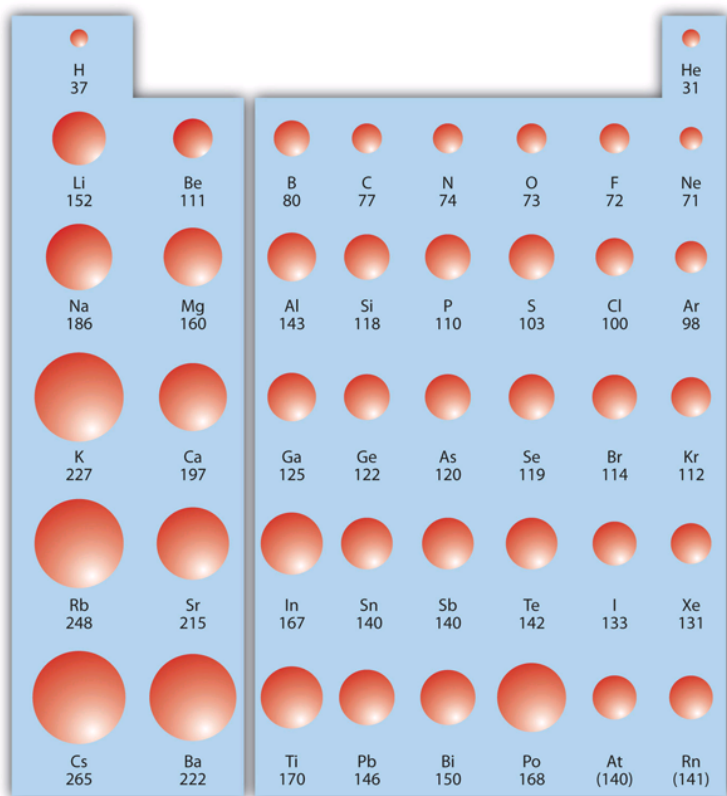


Figure 2.4.1: Atomic Radii trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in picometres. Image attribution: “Atomic Radii Trends on the Periodic Table” by David W. Ball © CC BY-NC-SA (Attribution-NonCommercial-ShareAlike).

Atoms as Balls: Why the approximation?

While atoms aren't small balls floating around in space, it can be handy for us to treat them as such! In constructing 3D models of compounds, the relative sizes of molecules should be taken into consideration. You don't need to know the specifics, but should be able to take an educated guess at what is bigger on the periodic table. Consider H_2O , how big are the hydrogen atoms compared to the oxygen?

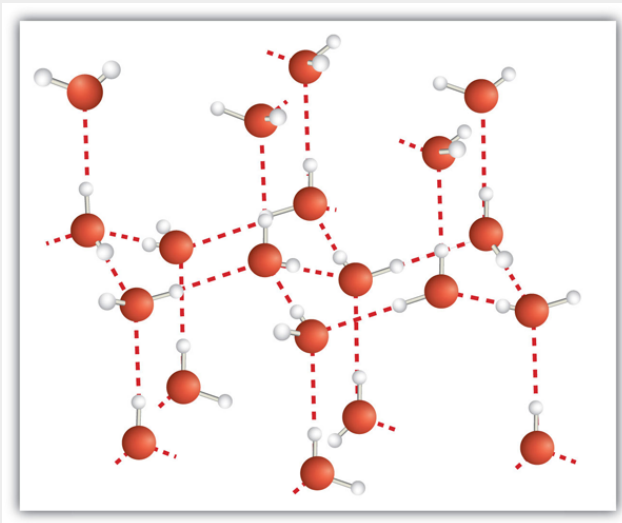


Figure 2.4.2: Water as a solid (ice). Note the size difference between the (white) hydrogen atoms and the (red) oxygens. With the radius roughly doubled — the area of the circle should increase by a factor of 4. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

In chapter 7.2, Thermochemistry Essentials, we will talk about collision theory — in which atoms and molecules are seen as whole spherical objects that can collide with one another. In tertiary studies, we can mathematically work out the chances of a collision occurring through the size of the cross-section of a molecule. It is much easier to determine the probability of a sphere colliding than a complex object. For simple

compounds, such as **diatomic molecules** comprised of a large atom and a relatively small atom, this approximation works.

Naturally, simplifications like this always bring about inaccuracy in our answers — however these are acceptable at this stage in your chemical career!

Ionisation energy

The first **ionisation energy** is the minimum energy required to remove one electron from a neutral atom. The second and third ionisation energies are the quantities necessary to remove the second and third electron from the atom, respectively. Ionisation energy increases when moving from left to right across a period as electrons are bound tightly. When proceeding down a group, the first ionisation energy decreases as an electron in a higher energy level is easier to remove. The trend in the first ionisation energy is the inverse of the atomic radii (see Figure 2.4.3). As atomic radii increase, ionisation energy decreases.

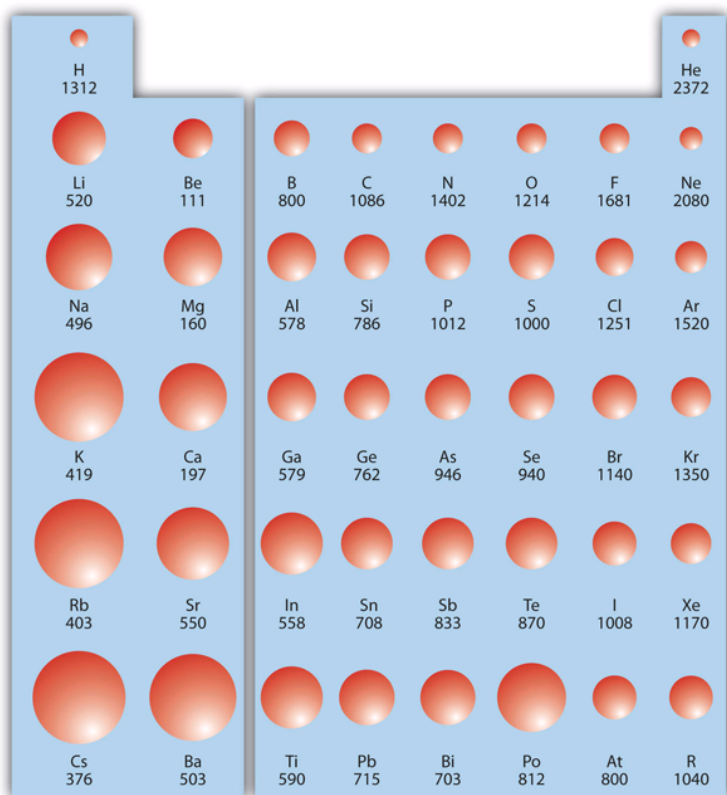


Figure 2.4.3 “Ionisation Energy on the Periodic Table.” Values are in kJ/mol. View accessible transcript here. Image attribution: “Ionization Energy on the Periodic Table” by David W. Ball © CC BY-NC-SA (Attribution-NonCommercial-ShareAlike).

Electronegativity

Electronegativity is the power of an atom in a molecule to attract electrons. The larger the value, the larger the electron-attracting ability. Atoms with higher electronegativity form **anions**, whereas atoms with smaller electronegativity form **cations** (see Figure 2.4.4).

Electronegativity decreases from top to bottom and increases from the left to the right of the periodic table – similar to the trend of ionisation energy (see Figure 2.4.3). Electronegativity is an important elemental property, as it dictates the types of bonds that can form between elements.

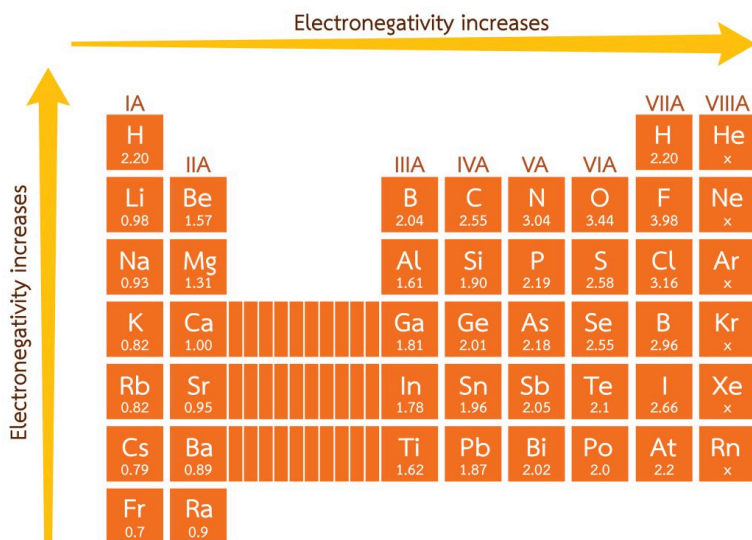


Figure 2.4.4: Electronegativity of elements. Image attribution: Periodic table of elements © Torsu – stock.adobe.com.

Key Takeaways

- Due to the construction of the periodic table, a number of trends can be seen moving across it.
- Atomic radius of atoms increases from the top to

the bottom of a group, and gets larger as one moves from the right to the left.

- Ionisation energy, the energy needed to remove an electron, decreases as atomic radius increases.
- Electronegativity, the power of an atom to attract electrons within molecules, increases alongside ionisation energy

Exercises

Practice Questions



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Transcript

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2.5 Isotopes and Relative Abundance

Learning Objectives

1. Understand what an isotope is and how they differ from normal atoms.
2. Express the masses of atoms with reference to relative abundance.
3. Apply atomic masses to the mass of molecules.

Because matter is defined as anything that has mass and takes up space, it should not be surprising to learn that atoms and molecules have mass.

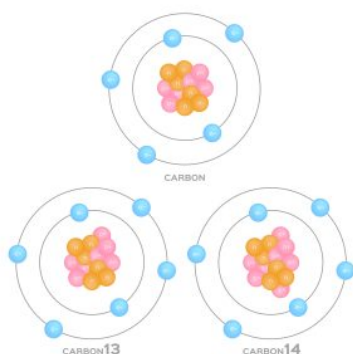


Figure 2.5.1: Stable Isotopes of Carbon.
Image attribution: isotope of carbon . 3
steps of carbon . vector © gritsalak –
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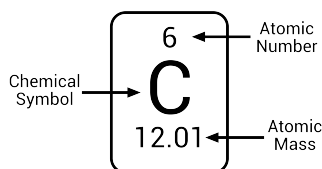
So far, we have assumed that all elements consist of the same mass. However, elements exist in a variety of **isotopes** – elements with the same amount of **protons**, but varying numbers of **neutrons**. Isotopes are named according to their element and **atomic mass**.

Individual atoms and molecules, however, are very small, and the masses of individual atoms and molecules are also very small. For macroscopic objects, we use

units such as grams and kilograms to state their masses, but these units are much too big to comfortably describe the masses of individual atoms and molecules. As the mass of atoms is incredibly tiny (in the realm of $0.000000000000000000000001g$ or $1 \times 10^{-24}g$), the **Dalton (Da)** or **atomic mass unit (amu)** is used as a reference (both can be used interchangeably). Protons and neutrons weigh 1 amu each (the mass of electrons is negligible).

For instance, the 6 proton carbon atom has three stable, naturally occurring isotopes: carbon-12 ($^{12}_6\text{C}$), carbon-13 ($^{13}_6\text{C}$) and carbon-14 ($^{14}_6\text{C}$), weighing 12, 13 and 14 amu respectively. This extra weight is accounted for by the additional neutrons added. In chapter 5, we will discuss how **Avogadro's number** connects Daltons and atomic mass units to a more applicable measurement: grams per mole ($g\ mol^{-1}$).

Atomic mass and relative percentage abundance



Almost all elements exist as a mixture of isotopes. The atomic mass of an element is a weighted average of the masses of the isotopes that compose an element. A weighted average is found by multiplying each mass

Figure 2.5.2: Periodic table entry of carbon. Note how the atomic mass reflects the fact that Carbon-12 is the most abundant isotope of carbon. Image attribution: parts of element on periodic table. carbon parts. chemical symbol atomic symbol and atomic mass © SAMYA – stock.adobe.com.

by its fractional occurrence (or relative abundance) and then adding all the products. The sum is the weighted average and serves as the formal atomic mass of the element. As such, the atomic masses listed on the periodic table take into account the **relative abundance** of each isotope. For instance, of all the carbon present, carbon-12 makes up 98.9% of it, while around 1% exists as carbon-13 and 1×10^{-10} exists as carbon-14¹. To determine the atomic mass to list on the periodic table, we take the atomic masses of these isotopes and adjust them by their relative abundance, as seen in Table 2.5.1.

Isotopes

Carbon-12

Carbon-13

Carbon-14

1. cited from W.M. Haynes CRC Handbook of Chemistry 97th Edu.

Carbon's atomic mass is calculated by multiplying the atomic mass of each isotope by its percentage abundance and then totalling the answers obtained for each isotope.

Carbon-12:

$$\frac{98.9}{100} \times 12.00 \text{ amu} = 11.868 \text{ amu}$$

Carbon-13:

$$\frac{1.1}{100} \times 13.00 \text{ amu} = 0.143 \text{ amu}$$

Carbon-14:

$$\frac{1 \times 10^{-10}}{100} \times 14.00 \text{ amu} = 1.4 \times 10^{-11} \text{ amu}$$

The atomic mass of carbon
 $= 11.868 \text{ amu} + 0.143 \text{ amu} + 1.4 \times 10^{-11} \text{ amu} = 12.011 \approx 12.01 \text{ amu}$

Taking the weighted averages of masses, the atomic mass displayed for carbon on the periodic table is around 12.01 amu.

Virtually all elements exist as mixtures of isotopes, so atomic masses may vary significantly from whole numbers. Table 6.1 "Selected Atomic Masses of Some Elements" lists the atomic masses of some elements. The atomic masses in Table 6.1 are listed to three decimal places where possible, but in most cases, only one or two decimal places are needed. Note that many of the atomic masses, especially the larger ones, are not very close to whole numbers. This is, in part, the effect of an increasing number of isotopes as the atoms increase in size. (The record number is 10 isotopes for tin.)

Table 2.5.1 Selected Atomic Masses of Some Elements

Element name	Atomic mass (amu)
Aluminum	26.981
Argon	39.948
Arsenic	74.922
Barium	137.327
Beryllium	9.012
Bismuth	208.980
Boron	10.811
Bromine	79.904
Calcium	40.078
Carbon	12.011
Chlorine	35.453
Cobalt	58.933
Copper	63.546
Fluorine	18.998
Gallium	69.723
Germanium	72.64
Gold	196.967
Helium	4.003
Hydrogen	1.008
Iodine	126.904
Iridium	192.217
Iron	55.845
Krypton	83.798
Lead	207.2
Lithium	6.941
Magnesium	24.305
Manganese	54.938
Mercury	200.59

Element name	Atomic mass (amu)
Molybdenum	95.94
Neon	20.180
Nickel	58.693
Nitrogen	14.007
Oxygen	15.999
Palladium	106.42
Phosphorus	30.974
Platinum	195.084
Potassium	39.098
Rubidium	85.468
Scandium	44.956
Selenium	78.96
Silicon	28.086
Silver	107.868
Sodium	22.990
Strontium	87.62
Sulphur	32.065
Tantalum	180.948
Tin	118.710
Titanium	47.867
Tungsten	183.84
Uranium	238.029
Xenon	131.293
Zinc	65.409
Zirconium	91.224

Now that we understand that atoms have mass, it is easy to extend the concept to the mass of molecules. The molecular mass is the sum of the masses of the atoms in a molecule. This may seem like a trivial extension of the concept, but it is important to count the number of each type of atom in the molecular formula. Also, although each atom in a molecule is a particular isotope, we use the weighted average, or atomic mass, for each atom in the molecule.

For example, if we were to determine the molecular mass of dinitrogen trioxide, N_2O_3 , we would need to add the atomic mass of nitrogen two times with the atomic mass of oxygen three times:

$$\begin{array}{rclcl}
 2 \text{ N masses} & = & 2 \times 14.007 \text{ amu} & = & 28.014 \text{ amu} \\
 + \quad 3 \text{ O masses} & = & 3 \times 15.999 \text{ amu} & = & 47.997 \text{ amu} \\
 \hline
 & & \text{total} & = & 76.011 \text{ amu} = \text{the molecular mass of } \text{N}_2\text{O}_3
 \end{array}$$

We would not be far off if we limited our numbers to one or even two decimal places.

Chemistry Is Everywhere: Sulphur Hexafluoride

On March 20, 1995, the Japanese terrorist group Aum Shinrikyo (Sanskrit for “Supreme Truth”) released some sarin gas in the Tokyo subway system; twelve people were killed, and thousands were injured. Sarin (molecular formula $\text{C}_4\text{H}_{10}\text{FPO}_2$) is a nerve toxin that was first synthesised in 1938. It is regarded as one of the most deadly toxins known, estimated to be about 500 times more potent than cyanide. Scientists and engineers who study the spread of chemical weapons such as sarin would like to have a less dangerous

chemical, ideally one that is nontoxic, so they are not at risk themselves.

Sulphur hexafluoride is used as a model compound for sarin. SF_6 has a similar molecular mass (about 146 u) as sarin (about 140 u), so it has similar physical properties in the vapour phase. Sulphur hexafluoride is also very easy to accurately detect, even at low levels, and it is not a normal part of the atmosphere, so there is little potential for contamination from natural sources. Consequently, SF_6 is also used as an aerial tracer for ventilation systems in buildings. It is nontoxic and very chemically inert, so workers do not have to take special precautions other than watching for asphyxiation.

Sulphur hexafluoride also has another interesting use: a spark suppressant in high-voltage electrical equipment. High-pressure SF_6 gas is used in place of older oils that may have contaminants that are environmentally unfriendly. Additionally, being much denser than air, it can give a person a much lower voice – the opposite to helium! Compounds often have a variety of uses, with more always being discovered from the scientific community.



One or more interactive elements has been excluded from this version of the text. You can view them online here:

[https://rmit.pressbooks.pub/
rmitchemistrybridgingcourse/?p=684#oembed-1](https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=684#oembed-1)

Key Takeaways

- Elements have a variety of isotopes: atoms with the same number of protons but different numbers of neutrons.
- Isotopes are named by their element and the number of subatomic particles in the nucleus. (Such as carbon-12, carbon-13, carbon-14).
- Atomic masses found on the periodic table are a weighted average of the relative abundance of all isotopes.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

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- parts of element on periodic table. carbon parts. chemical symbol atomic symbol and atomic mass © SAMYA - stock.adobe.com is licensed under a All Rights Reserved license

PART III

CHAPTER 3: CHEMICAL BONDS

3.1 The Octet Rule and Lewis Dot Diagrams

Learning Objectives

1. Understand the octet rule and how to read and write Lewis dot diagrams
2. Utilise Lewis dot diagrams to predict the bonding of simple molecules.

In 2.2 ‘Electronic Configuration’, we discussed how **valence electrons** are essential to determining the bonding of elements. Generally speaking, elements want to achieve 8 electrons in their outer shell – the same configuration seen in noble gases – and will bond with other elements to make this occur. This rule of thumb is known as the **octet rule**, and is true for elements up to the *d* block.

The Lewis electron-dot model is based on the fact that atoms tend to combine to give compounds that achieve a noble gas valence electron configuration of s^2p^6 (or s^2 for hydrogen). Elements within this model are represented through their chemical symbol and dots, which depict valent *s* and *p* electrons. Following the rules of electronic orbitals, electrons are placed surrounding each of the four sides of the element before paring up (see Figure 3.1.1).

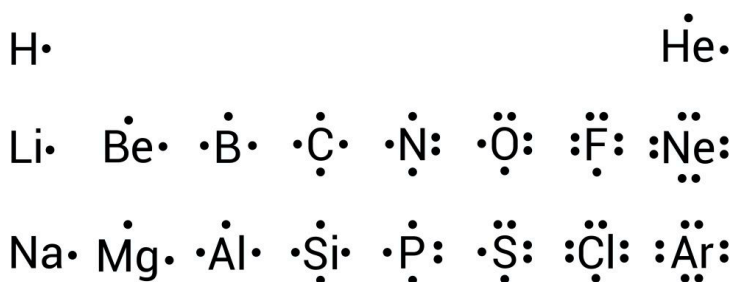


Figure 3.1.1: Lewis Electron Dot Diagrams. Image attribution: Lewis dot structure of elements 1 to 18. Periodic table © Zizo – stock.adobe.com.
Transcript.

To understand how dot diagrams can help us predict bonding, let's have a look at a compound containing sodium (**Na**) and chlorine (**Cl**). We will begin by determining the number of valence shell electrons in each:

Table 3.1.1: Electron configurations of sodium and chlorine.

Atom	Atomic number	Electron configuration	Valence shell electrons
Na	11	$1s^2 2s^2 2p^6 3s^1$	1
Cl	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	7

Sodium contains 1 valence electron, while chlorine has 7. For both to achieve full valence shells, sodium can donate its electron to chlorine.

The atom that can form most bonds is considered the central atom. If two different atoms can form the same number of bonds, then the least electronegative atom should be the central atom. Among sodium and chlorine, sodium is the least electronegative

atom. Let us draw the two elements as Lewis Dot Diagrams (see Figure 3.1.2):

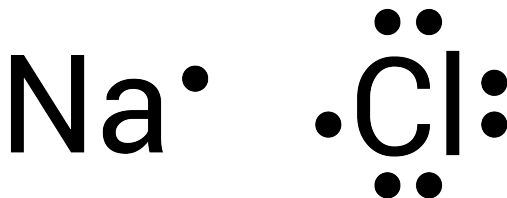
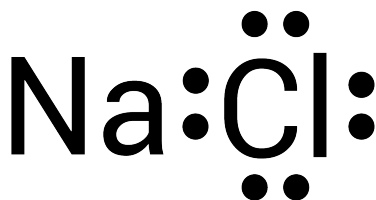


Figure 3.1.2: Sodium and chlorine in Lewis dot diagram.

To create this neutral compound, a single bond is formed between sodium and chlorine to donate electrons. This is depicted by combining the two dots together or by drawing a line (see Figure 3.1.3):



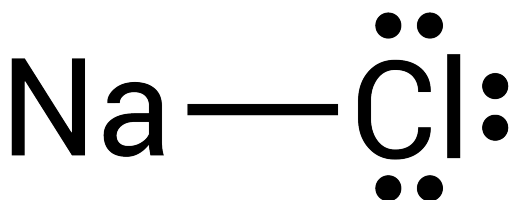


Figure 3.1.3: Sodium and chlorine form a bond by sharing an electron in a Lewis dot diagram. The bond can be shown as a line or two electrons in Lewis dot diagrams.

Chemistry Is Everywhere: Salt

The element sodium is a very reactive metal; given the opportunity, it will react with the sweat on your hands and form sodium hydroxide, which is a very corrosive substance. The element chlorine is a pale yellow, corrosive gas that should not be inhaled due to its poisonous nature. Bring these two hazardous substances together, however, and they react to make the ionic compound sodium chloride, known simply as salt. Sodium, chlorine, and sodium chloride can be seen in Figure 3.1.4 below:

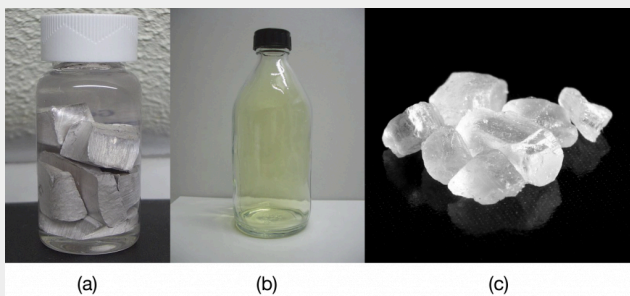


Figure 3.14 “Sodium Chloride.” (a) Sodium is a very reactive metal. (b) Chlorine is a pale yellow, noxious gas. (c) Together, sodium and chlorine make sodium chloride – salt – which is necessary for our survival. Image attribution: “Sodium metal chunks in oil” by Wilco Oelen © CC BY-SA (Attribution-ShareAlike). “Chlorine in bottle” by Wilco Oelen © CC BY-SA (Attribution-ShareAlike). “Salt Crystals” by Mark Schellhase © CC BY-SA (Attribution-ShareAlike).

Salt is necessary for life. Na^+ ions are one of the main ions in the human body and are necessary to regulate the fluid balance in the body. Cl^- ions are necessary for proper nerve function and respiration. Both of these ions are supplied by salt. The taste of salt is one of the fundamental tastes; salt is probably the most ancient flavouring known and one of the few rocks we eat.

The health effects of too much salt are still under debate, although a 2010 report by the US Department of Agriculture concluded that “excessive sodium intake ... raises blood pressure, a well-accepted and extraordinarily common risk factor for stroke, coronary

heart disease, and kidney disease.”¹ It is clear that most people ingest more salt than their bodies need, and most nutritionists recommend curbing salt intake. Curiously, people who suffer from low salt (called *hyponatremia*) do so not because they ingest too little salt but because they drink too much water. Endurance athletes and others involved in extended strenuous exercise need to watch their water intake so their body’s salt content is not diluted to dangerous levels.

This completes the diagram for sodium chloride. Let’s have a look at a more complex molecule, ammonia (NH₃). Ammonia contains nitrogen and hydrogen. As before, we will begin by considering the valence electrons.

Table 3.1.2: Electron configuration of nitrogen and hydrogen.

Atom	Atomic number	Electron configuration	Valence shell electrons
H	1	$1s^1$	1
N	7	$1s^2 2s^2 2p^3$	5

Hydrogen contains 1 valence electron, while nitrogen has 5. To

1. U.S. Department of Agriculture Committee for Nutrition Policy and Promotion, “Report of the Dietary Guidelines Advisory Committee on the Dietary Guidelines for Americans,” accessed January 5, 2010, <https://www.dietaryguidelines.gov/sites/default/files/2019-05/2010DGACReport-camera-ready-Jan11-11.pdf>.

achieve a full valence shell, 3 hydrogens are needed to donate their electrons with nitrogen (see Figure 3.1.5).

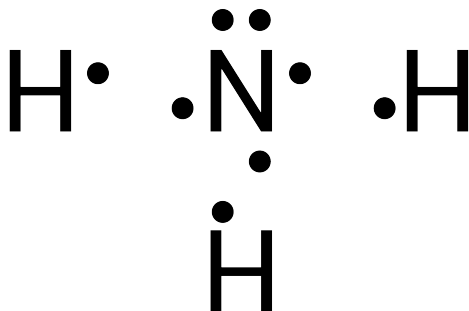


Figure 3.1.5: Nitrogen and hydrogen in Lewis dot diagram.

To create this neutral compound, a single bond is formed between each of the hydrogen and the lone nitrogen to donate electrons (see Figure 3.1.6):

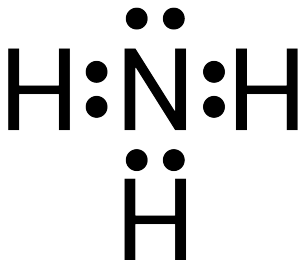


Figure 3.1.6: Nitrogen and hydrogen in Lewis dot diagram, forming bonds through sharing electrons.

Lewis dot diagrams are vital in understanding the geometry of molecules. As we will explore in VSEPR theory – the location of lone pairs of electrons are of great importance. Let us review one more example: carbon dioxide (CO₂).

Table 3.1.3: Electron configuration of carbon and oxygen.

Atom	Atomic number	Electron configuration	Valence shell electrons
C	6	$1s^2 2s^2 2p^2$	4
O	8	$1s^2 2s^2 2p^4$	6

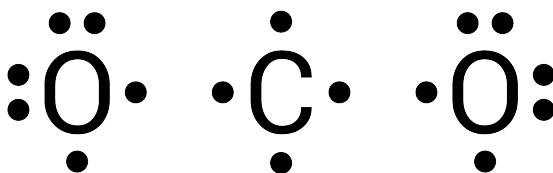


Figure 3.1.7: Carbon and oxygen in Lewis dot diagram.

In this compound, carbon is the central atom (see Figure 3.1.7). Carbon requires 4 more electrons to complete its valence shell. We can begin by donating one electron from each neighbouring oxygen atom (see Figure 3.1.8):

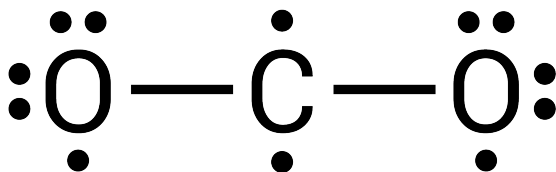


Figure 3.1.8: Carbon and oxygen in Lewis dot diagram, forming a bond through sharing electrons.

After completing this, we can see how all atoms still have unpaired electrons present. Two on the carbon and one on each of the oxygens. For a stable compound to form, all electrons must be paired. This, therefore, must mean that the oxygens are **double bonded** to the carbon, where two pairs of electrons are transferred/shared between elements (see Figure 3.1.9). Within chemistry, **triple bonds** are also a possibility.

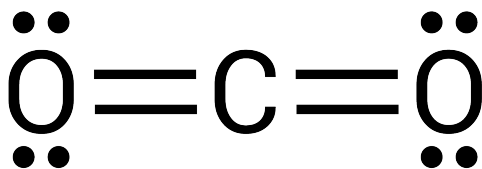


Figure 3.1.9: Carbon and oxygen in Lewis dot diagram, forming a double bond through sharing additional electrons.

Some molecules do not follow the octet rule. Atoms such as sulphur, phosphorus and chlorine deviate from the octet rule by having more than eight electrons in the valence shell. This is due to the availability of vacant *d* orbitals in the third shell. These exceptions are fully explored in tertiary chemistry studies.

Key Takeaways

- Main group elements want to achieve 8 electrons in their outer shell. This is known as the octet rule.
- Lewis electron-dot diagrams allow for us to visually represent valent electrons and predict bonding.
- Some molecules require the formation of double and triple bonds to complete their octet.
- Not all compounds follow the traditional octet rule, such as sulphur and phosphorous.

Exercises

Practice Questions



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3.2 Ionic Bonds and Electronegativity

Learning Objectives

1. Know how ions form.
2. Learn the characteristic charges that ions have.
3. Construct a proper formula for an ionic compound.
4. Generate a proper name for an ionic compound.

So far, we have discussed elements and compounds that are electrically neutral. However, this is not always the case. As we have seen within Lewis Dot Diagrams, electrons can move from one atom to another; when they do, **species** with overall electric charges are formed. Such species are called **ions**. Species with overall positive charges are termed **cations**, while species with overall negative charges are called **anions** (seen in Table 3.2.1). Remember, ions are formed only when *electrons* move from one atom to another; a proton never moves from one atom to another. Compounds formed from positive and negative ions are called ionic compounds.

Table 3.2.1 Monatomic Ions of Various Charges.

Ions formed by losing a single electron	H ⁺	Ions formed by gaining a single electron	F ⁻
	Na ⁺		Cl ⁻
	K ⁺		Br ⁻
	Rb ⁺		I ⁻
	Ag ⁺		
	Au ⁺		
Ions formed by losing two electrons	Mg ²⁺	Ions formed by gaining two electrons	O ²⁻
	Ca ²⁺		S ²⁻
	Sr ²⁺		Se ²⁻
	Fe ²⁺		
	Co ²⁺		
	Ni ²⁺		
	Cu ²⁺		
	Zn ²⁺		
	Sn ²⁺		
	Hg ²⁺		
	Pb ²⁺		
Ions formed by losing three electrons	Sc ³⁺	Ions formed by gaining three electrons	N ³⁻
	Fe ³⁺		P ³⁻
	Co ³⁺		
	Ni ³⁺		
	Au ³⁺		
	Al ³⁺		
	Cr ³⁺		
Ions formed by losing four electrons	Ti ⁴⁺		
	Sn ⁴⁺		
	Pb ⁴⁺		

As noted in the table, some elements will appear multiple times. For instance, all metals can form more than one possible charge. For example, iron atoms can form 2+ cations or 3+ cations. Cobalt is another element that can form more than one possible charged ion (2+ and 3+), while lead can form 2+ or 4+ cations. Unfortunately, there is little understanding of which two charges a metal atom may take, so it is best to just memorise the possible charges a particular element can have.

Note the convention for indicating an ion. The magnitude of the charge is listed as a right superscript next to the symbol of the

element. If the charge is a single positive or negative one, the number 1 is not written; if the magnitude of the charge is greater than 1, then the number is written *before* the + or – sign. An element symbol without a charge written next to it is assumed to be the uncharged atom.

Naming Ions

Main group cations are named using the name of the element followed by ion. For instance, Li^+ : Lithium ion, Mg^{2+} : Magnesium ion and Al^{3+} : Aluminium ion.

Transition metals (*d*, *f* block) and some *p* block elements can form multiple cations such as Cu^+ and Cu^{2+} , Pb^{2+} and Pb^{4+} . These types of cations can be named in two ways. The first is that the ion with the smaller charge ends with “-ous” and the larger charge ends with “-ic”. For instance, Cu^+ is cuprous ion and Cu^{2+} is cupric ion. The more common way is writing the charge of the ion in Roman numerals within parenthesis next to the element name. For example, Cu^+ is Copper(I) ion, and Cu^{2+} is Copper(II) ion.

Anions are named by removing the ending of the element name and replacing it with “-ide”. For instance, the anion formed by Cl is known as a chloride ion.

Polyatomic ions consist of more than one atom. For example, SO_4^{2-} (sulphate), NO_3^- (nitrate), PO_4^{3-} (phosphate) and NH_4^+ (ammonium) ions. Atoms in polyatomic ions are held together by **covalent bonds**. These generally have unique names as seen in Table 3.2.3.

Name

ammonium

acetate

bicarbonate (hydrogen carbonate)

bisulphate (hydrogen sulphate)

carbonate

chlorate

chromate

cyanide

dichromate

hydroxide

nitrate

nitrite

peroxide

perchlorate

permanganate

phosphate

sulphate

sulphite

triiodide

Ionic bonds

Ionic bonds are formed by the complete transfer of electrons from one atom to another. Ionic bonds are generally created between metals and nonmetals. Metals provide the cation (lose electrons), and nonmetals (gain electrons) provide the anion. For an ionic bond to form, there must be a large difference in electronegativity between two elements.



Figure 3.2.1 “NaCl = Table Salt.” The ionic compound NaCl is very common. Image attribution: “Kosher Salt” by stlbites.com © CC BY-ND (Attribution NoDerivs)

Electronegativity is the power of an atom in a molecule to attract electrons. If you recall from 2.4 Periodic Trends, electronegativity becomes higher as one moves from the bottom left of the periodic table to the halogens. As a result, bonds between elements from the far left and far right of the periodic table are typically ionic. For an ionic compound to be formed, the difference between the electronegativities of two participating atoms **must be larger than or equal to two**.

Please note that, unlike other measurements, electronegativity is a unitless quantity. It is not measured in a unit such as grams or metres — it is simply a number. See Figure 3.2.1 for the electronegativities of the main group elements.

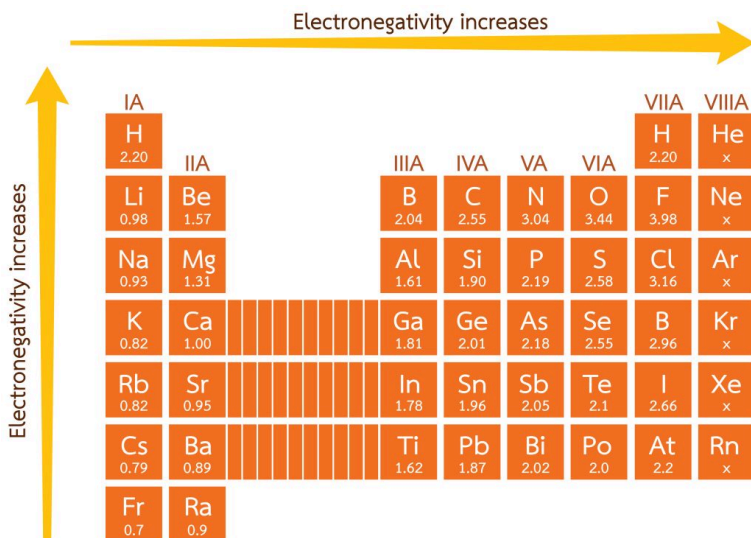


Figure 3.2.1: Electronegativity of Elements. Image attribution: Periodic table of elements © Torsu – stock.adobe.com. Transcript.

As an example, for the commonly-known ionic compound, sodium chloride (NaCl), the difference in electronegativity of **Na** ($\text{EN} = 0.93$) and **Cl** ($\text{EN} = 3.16$) is **2.23**. The larger difference in electronegativity creates electrostatic interactions between the electrons of one atom and the nucleus of the other atom. Ionic compounds are held together by electrostatic interactions between the cations and anions involved.

Names and Formulas for Ionic Compounds

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula has a **cation** and an **anion** in it; an ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is to ensure the total

positive charge balances out the total negative charge. Consider the ionic compound between Na^+ and Cl^- . Each ion has a single charge (one positive and one negative), so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions: (1) write the formula for the cation first and the formula for the anion next, but (2) do not write the charges on the ions. Thus, for the compound between Na^+ and Cl^- , we have the ionic formula NaCl .

Because the charges on the ions are constant, sometimes we have to have more than one cation or an anion to balance the overall positive and negative charges. For the ionic compound between Mg^{2+} ions and Cl^- ions, we now consider the fact that the

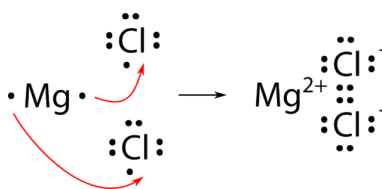


Figure 3.2.2: Formation of Magnesium Chloride. Image attribution: an image from CK12 Foundation CC BY-NC 3.0. Transcript.

charges have different magnitudes; $2+$ on the magnesium ion and $1-$ on the chloride ion. To balance the charges with the lowest number of ions possible, we need to have two chloride ions to balance the charge on the one magnesium ion. Rather than write the formula MgClCl , we combine two chloride ions and write it with a 2 subscript: MgCl_2 .

Naming ionic compounds requires you to combine the name of the cation and the name of the anion, in both cases omitting the word *ion*. Do not use numerical prefixes if more than one ion is necessary to balance the charges. NaCl is sodium chloride, a combination of the name of the cation (sodium) and the anion (chloride). MgO is magnesium oxide. MgCl_2 is magnesium chloride – *not* magnesium dichloride.

In naming ionic compounds whose cations can have more than one possible charge, we must also include the charge, in parentheses and in roman numerals, as part of the name. Hence FeS is iron(II) sulphide, while Fe_2S_3 is iron(III) sulphide. No

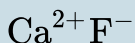
numerical prefixes appear in the name. The number of ions in the formula is dictated by the need to balance the positive and negative charges.

Example: Determine the chemical formula of the compound produced when Ca^{2+} and F^{-} ions interact.

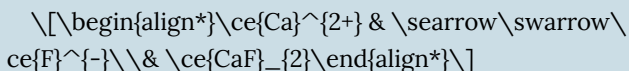
There are two ways of writing chemical formulas.

Method 1:

First, write down the symbol of the cation with its charge, followed by the anion.



Next, cross each ion's charge (only the numerical value, not the sign) and write these numbers down as subscripts. So, the calcium ion's charge will be written next to the fluoride ion, and the fluoride ion's charge will be written next to the calcium ion, in both cases as subscripts:



Method 2:

Ca ion has a **+2** charge, and **F** ion has a **−1** charge. Two **F** ions, each with **−1** charge, will neutralise the **+2** charge of **Ca** ion. Therefore, the chemical formula is **CaF₂**

Physical Properties of Ionic Compounds

Your favourite energy sports drinks (see Figure 3.2.3) typically love to advertise the number of electrolytes they have; but what exactly are they? When one exercises or is sick, the sweat excreted causes the body to lose essential ions important for carrying electrical signals. Ions such as Na^+ , K^+ and Cl^- have to be continuously replaced.



Figure 3.2.3: Athletes love sports drinks, but what is in them that's so important? Image attribution: Sporty runners couple drinking energy drink © Davide Angelini – stock.adobe.com.

Sports drinks and rehydration solutions contain increased amounts of these ions to ensure one retains a healthy number of conductive ions.

Ionic compounds are typically crystalline solids (like table salt) with high melting and boiling points that are soluble in **polar** solvents (such as water). When in their **aqueous** state, ionic compounds are able to conduct electricity through the movement of ions.

Key Takeaways

- Ions form when species lose or gain electrons, commonly when compounds dissolve in water.

- Metal ions have a variety of oxidation states, and can form numerous different ions.
- Cations refer to positively charged ions, while anions refer to negatively charged.
- When the difference in electronegativity is 2 or more, an ionic bond forms between two atoms.
- The main group cations are named using the name of the element followed by an ion, while anions feature the suffix “-ide.”
- Many polyatomic ions exist that are commonly referred to as ammonium, cyanide, and hydroxide.
- Ions in solution are conductors of electricity, an important property for healthy living within living beings.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

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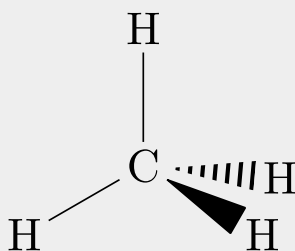
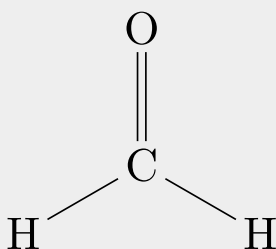
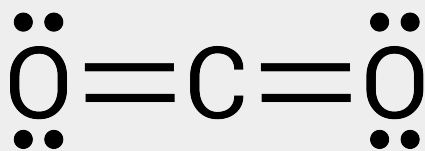
3.3 Covalent Bonds and Polarity

Learning Objectives

- Describe a non-polar bond and a polar bond.
- Use electronegativity to determine whether a bond between two elements will be nonpolar covalent, polar covalent, or ionic.

While **ionic bonds** form when metals **donate electrons** to non-metals, **covalent bonds** are formed between two non-metals **sharing electrons**. As a result of this, covalent bonds form when the difference in **electronegativity** between elements is below 2. Covalent bonds are the strongest kind of bond that can be formed. Atoms held together by covalent interactions give rise to molecular compounds.

Similar to ionic compounds, atoms involved in a covalent bond acquire the octet in their valence shell by sharing electrons between atoms. Generally, the number of covalent bonds an atom forms depends on how many electrons an atom requires to achieve an octet. For instance, C forms four covalent bonds as the electronic configuration ($1s^2 2s^2 2p^2$) needs four more electrons to acquire an octet in the valence shell.



Carbon's ability to form 4 covalent bonds makes it a vital element – with some regard it as the most important element in life. Carbon's prevalence within

many molecules is detailed in the subject of organic chemistry — which explores the synthesis of a variety of compounds and their functional groups.

Polarity

Covalent bonds are formed through the sharing of electrons. The difference in electronegativity is still, however, important, as it indicates if the shared electron can be found closer to one atom or another. Let's observe how this presents in hydrogen fluoride (see Figure 3.3.1):

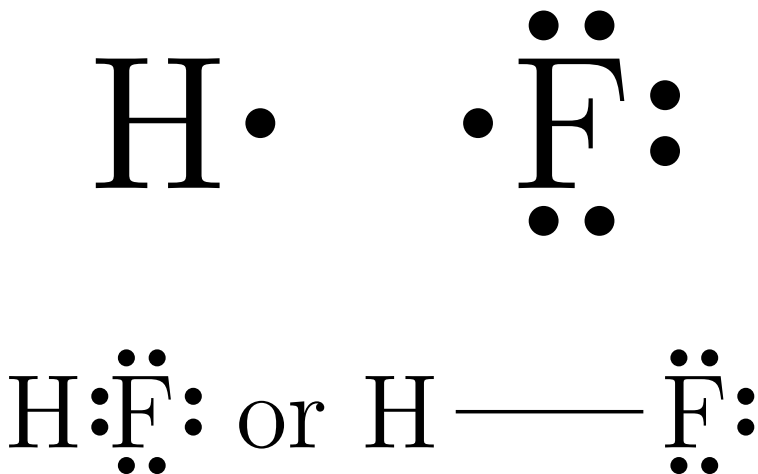


Figure 3.3.1: Lewis dot diagram and bonding of hydrogen fluoride.

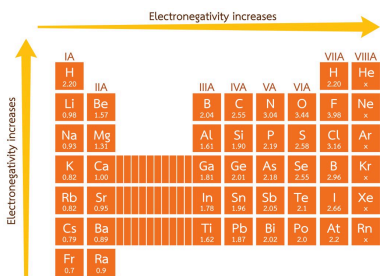


Figure 3.3.2: Electronegativity of Elements. Image attribution: Periodic table of elements © Torsu – stock.adobe.com. Transcript.

Utilising the electronegativity values presented in Figure 3.3.2, **H** is valued at 2.20, where **F** presents at 3.98. The difference in electronegativity is 1.78. While this isn't enough to qualify as an ionic bond, the higher electronegativity on the fluoride atom causes the electron to be positioned in space closer towards it. This

generates a **partial charge**, represented by a lowercase delta (δ). The more electronegative atom gains a partial negative charge due to the electron being closer to the atom, while the less electronegative generates a partial positive (see Figure 3.3.3). The presence of these partial charges gives rise to a very important chemical concept: **polarity**. These types of bonds, such as the one in hydrogen fluoride, are known as polar covalent bonds.

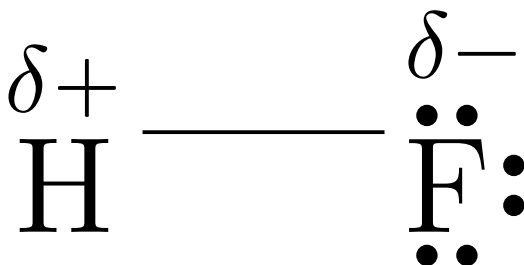


Figure 3.3.3: Partial charges on a hydrogen fluoride molecule.

The net polarity of a molecule depends on the polarity of the individual bonds, the contribution from lone pairs, and the molecules' shape. In a polar molecule, electrons are more strongly attracted to one part of the molecule than the other. For example, in the water molecule, due to the high electronegativity of the oxygen atom, electrons are more strongly attracted towards the

Note: All ionic bonds are polar by the nature of electronegativity .

oxygen atom than they are to the hydrogen atom. In some cases, although the individual covalent bonds are polar, the overall molecule becomes non-polar due to the shape of the molecule. For instance, the carbon dioxide molecule has polar covalent $\text{C} - \text{O}$ bonds. However, due to the linear symmetrical shape of the molecule, polar bonds (dipoles) cancel each other out, creating a zero net polarity (see Figure 3.3.4).

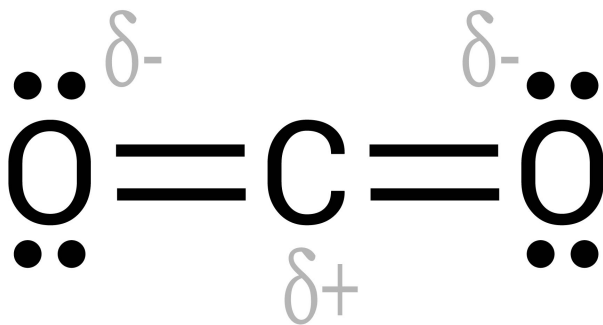


Figure 3.3.4: Carbon dioxide, although containing polar covalent bonds, is a non-polar molecule, as the electronegative ends directly oppose each other.

A lot of elements within nature that are gases at room temperature also form covalent bonds with themselves in order to complete their

octets. H_2 , N_2 and O_2 are some examples (see Figure 3.3.5). With the difference in electronegativity between the two atoms being 0 (as they are the same), the type of bond that forms is a non-polar covalent bond.

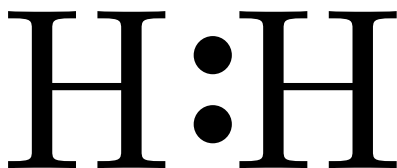


Figure 3.3.5: Hydrogen naturally exists as a diatomic molecule.

The net polarity of a molecule has implications regarding melting/boiling points, viscosity and, most importantly, solubility. We will discuss how to determine a molecule's polarity more during 3.6 Predicting Molecular Shape. For now, we can identify polar covalent from normal covalent bonds through the difference in electronegativity detailed in Table 3.3.1.

Table 3.1.1: Electronegativity Values for Bond Types.

Electronegativity difference	Type of bond	Examples
>2.0	Ionic	NaCl , CsF
0.5-1.9	Polar covalent	HCl , HF
0-0.4	(non-polar) Covalent	CH_4 , O_2 , H_2

To understand how the polarity of a bond differs with electronegativity, have a look at the following interactive activity's "Three Atoms" simulation with partial charges enabled:



One or more interactive elements has been excluded from this version of the text. You can view them online here: <https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=703>

Simulation by PhET Interactive Simulations, University of Colorado

Naming Binary Molecular Compounds

To name covalent compounds with 2 elements, take the name of the less electronegative atom first, followed by the more electronegative atom. Therefore, metals come before nonmetals and nonmetals located towards the left side of the periodic table. Additionally, the name of the more electronegative atom should end with the suffix *-ide* as used for anions. Depending on the number of atoms of each element involved in the compound, prefixes such as *di-*, *tri-*, *tetra-* and *penta-* are utilised (see Table 3.3.2). For instance, the compound CCl_4 is named as “carbon tetrachloride”.

Table 3.3.2: Named Examples of Covalent Compounds.

Molecular formula	Name of the compound
CO	Carbon monoxide
CO ₂	Carbon dioxide
BBr ₃	Boron tribromide
N ₂ O ₅	Dinitrogen pentoxide

Key Takeaways

- Covalent bonds form between two non-metals sharing electrons.
- Covalent bonds are identified as having a difference in electronegativity lower than 2.
- When the difference in electronegativity is between 0.4-2, a polar covalent bond forms. This bond has partial charges, with the δ^- assigned to the more electronegative element.
- Binary compounds are named by the less electronegative element first, followed by the other element involved. Prefixes are added to denote how many of each element are present.

Exercises

Practice Questions



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Transcript

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3.4 Metallic Bonds

Learning Objectives

- Define the properties of metallic bonds and delocalised electrons
- Understand the general characteristics of metals.

To complete our types of bonding, metallic bonds present when **metals** are bonded to other metals together in a **lattice**. Metallic bonds exhibit unique sets of properties within compounds, all due to a sea of delocalised electrons (see Figure 3.4.1).

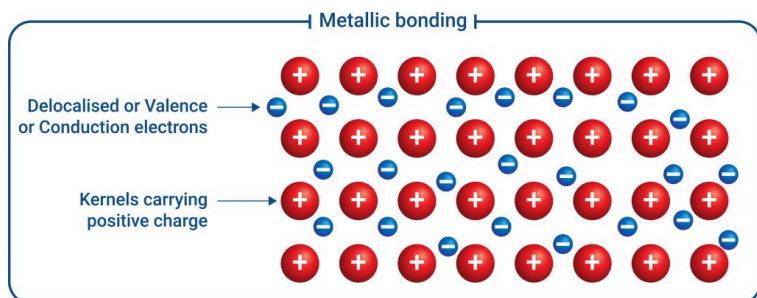


Figure 3.4.1: Diagram of Metallic Bonding. Image attribution: Metallic bonding: The electrostatic attractive force between the delocalised electrons present in the metallic lattice and the positively charged metal atom core. © Reuel Sa – stock.adobe.com.

In covalent and ionic bonds, the electron is tied between one or two atoms. Within metallic bonds, valence electrons are essentially liberated from any individual atom due to electron **orbitals** overlapping, allowing electrons to flow freely within the lattice. These delocalised electrons require significantly less energy to shift around due to a lack of attraction forces to the

nucleus. Even with just a small current, electrons will readily move throughout the structure. This makes metals fantastic conductors of electricity and heat!

The metallic lattice that forms and delocalised electrons cause metals to be **malleable** and **ductile**. Metals that are subjected to forces are able to shift their atoms whilst retaining the delocalised electrons and the lattice structure.

Metals are generally characterised as having high melting points and being lustrous, paramagnetic, and solid at room temperature; however, the individual electronic configuration of metallic elements does cause exceptions to arise. Mercury, for instance, is a transition metal but is a liquid at room temperature.

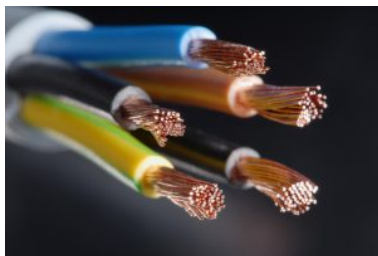


Figure 3.4.2: Copper wire is the backbone of a lot of electrical systems, thanks to copper's low-cost, ductility and ability to be drawn into a conducting wire. Image attribution: Kabel © demarco – stock.adobe.com.

You may notice we are glossing over metallic bonding! We stress that students are aware of the general characteristics of metals, being delocalised electrons leading to the high conduction of electricity and heat alongside high melting points due to the rigid lattice. You'll often be using hotplates or other metal implements in the lab, so understanding why they work is important! However, tertiary chemistry studies

don't focus on traditional metallic bonding. The study of inorganic chemistry focuses on the formation of metal complexes through bonding with 'ligands' that attract metal ions. Inorganic chemistry requires its own separate field of study, as metals can form many more bonds than would be traditionally expected (see Figure 3.4.3).

Inorganic chemistry often deals with how the oxidation states and coordination of transition metals change with respect to pH, heat and other variables. These oxidation states often produce energy changes with wavelengths within the visible spectrum, making inorganic chemistry one of the more colourful and

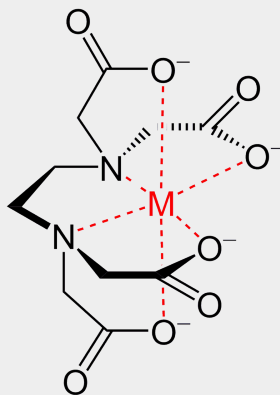


Figure 3.4.3: EDTA (Ethylenediaminetetraacetic acid) bonded to a metal ion. Image attribution: an image from the public domain courtesy of Wikimedia Commons.

exciting fields (see Figure 3.4.4)! For now, however, this study is outside the scope of this textbook.



Figure 3.4.3: Vanadium complexes of different oxidation states. Image attribution: an image from the public domain courtesy of Wikimedia Commons.

Key Takeaways

- Metallic bonding occurs when metals are bonded in a lattice.
- Metallic bonds feature a sea of delocalised electrons due to overlapping electron orbitals.
- Delocalised electrons allow for the easy conduction of heat and electricity compared to other elements and compounds.

Exercises

Practice questions



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Transcript

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- Metallic bonding: The electrostatic attractive force between the delocalised electrons present in the metallic lattice and the positively charged metal atom core. © Reuel Sa - stock.adobe.com
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3.5 Intermolecular Forces

Learning Objectives

- Understand the types of intermolecular forces: dipole-dipole interactions, hydrogen bonding and London dispersion forces, and their relative strengths.
- Learn how to identify what attraction forces are likely to present within a given molecule depending on its structure.

The bonds we have just discussed (covalent, ionic and metallic) are all examples of **intramolecular forces**: the forces inside a molecule that keep it together. However, molecules commonly interact with each other. For instance, the formation of solids is due to the compounds being within close proximity to each other. What is holding these molecules together?

Dipole-Dipole Interactions and Hydrogen Bonding

Like the ends of a magnet, opposite charges attract, while similar charges repel. As a result, attraction forces between molecules are naturally present in those with charge. Recall that hydrogen

chloride, a polar covalent molecule, experiences partial charges on its constituents. In this case, chlorine is more electronegative and is partially negative, given the closeness of the electron to the atom. On the other hand, hydrogen is partially positive because the electron is further away (see Figure 3.5.1).

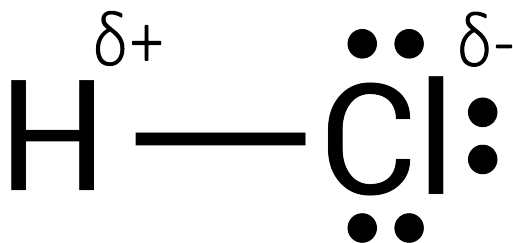


Figure 3.5.1: Hydrogen chloride (or more commonly hydrochloric acid) with partial charges.

When a molecule of hydrogen chloride is surrounded by another molecule of hydrogen chloride, it will begin to align according to attraction forces. Opposite ends will neighbour each other, as seen in Figure 3.5.2.



Figure 3.5.2: Hydrogen chloride molecules will begin to align positive and negative partial charges with each other, similar behaviour to that experienced by a magnet.

This is known as a **dipole-dipole interaction** – where molecules experience attraction and repulsion forces due to a permanent

charge. **Hydrogen bonds** are a special kind of dipole-dipole interaction, as hydrogen exhibits unusually strong intermolecular forces when paired with a highly electronegative atom such as chlorine or oxygen (see Figure 3.5.3).

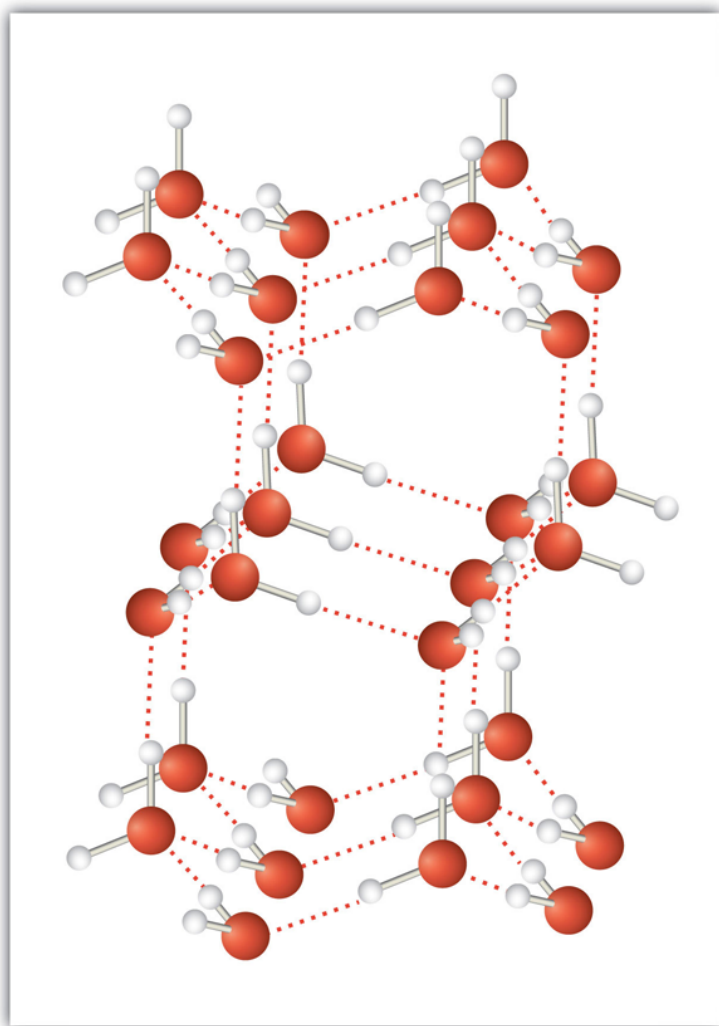


Figure 3.5.3: When water solidifies, hydrogen bonding between the molecules forces the molecules to line up in a way that creates empty space between the molecules, increasing the overall volume of the solid. This is why ice is less dense than liquid water. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

Intermolecular forces allow us to understand why molecules have a tendency to stay together (see Figure 3.5.3). It also explains the melting and boiling points of different compounds. Molecules that are closely attracted to one another require higher kinetic energy and, therefore, higher temperature to separate and undergo a phase change. The greater the strength of the

attraction and the more interactions that are allowed to occur, the greater the energy requirement. This is why ionic compounds (particularly salts) generally have high melting points. For reference, sodium chloride has a melting point of 802°C and a boiling point of 1465°C!

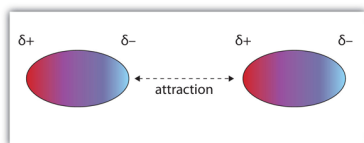


Figure 3.5.4: Dipole-Dipole Interactions. Oppositely charged ends of polar molecules attract each other. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

London Dispersion Forces

Under this theory, we should expect that for non-polar covalent molecules that don't hold permanent dipoles, melting and boiling points should be the same due to a lack of partial charge. However, as Table 3.5.1 demonstrates, boiling points do change with respect to the length of the molecule.

Table 3.5.1: Alcohol Table of Boiling Points (cited from W.M. Haynes CRC Handbook of Chemistry 97th Edu.).

Chemical Name	Molecular Formula	Boiling Point (°C)
Methane	CH_4	-161.5
Ethane	C_2H_6	-88.6
Propane	C_3H_8	-42.1
Butane	C_4H_{10}	-0.5
Pentane	C_5H_{12}	36.1
Hexane	C_6H_{14}	68.7
Heptane	C_7H_{16}	98.4
Octane	C_8H_{18}	125.6
Nonane	C_9H_{20}	150.8
Decane	$\text{C}_{10}\text{H}_{22}$	174.1

All molecules, regardless of dipoles, are subject to **dispersion forces** (also known as London forces). These occur due to electrons continuously shifting within molecules. These electrons may occasionally generate partial charges, which induce electrons in other atoms to do the same, generating a slightly attractive force between molecules. As the surface area of a molecule increases, so does the occurrence of dispersion forces. As a result, longer-chain alkanes (listed in Table 3.5.1) possess higher boiling points.

Those are the three kinds of **intermolecular forces**: hydrogen bonding, dipole-dipole interactions and London dispersion forces. You can see them listed in Table 3.5.2 and their relative strengths. In addition to dictating boiling and melting points within molecules,

they also play an important part in the solubility of molecules within water.

Table 3.5.2: Types of Intermolecular Forces.

Intermolecular Force	Type of Molecule	Strength of Attraction
Hydrogen Bonding	Molecules containing $H-X^1$	Very Strong
Dipole-Dipole Interactions	Polar Molecules	Strong
(London) Dispersion Forces	All molecules.	Weak

Key Takeaways

- Intramolecular forces are types of bonding (ionic, covalent and metallic), while intermolecular forces are the attraction and repulsion experienced between molecules.
- Dipole-dipole interactions are experienced between molecules with partial charges.
- Hydrogen bonding are a form of dipole-dipole interactions, but only feature on molecules with hydrogen next to a highly electronegative element.
- London dispersion forces are felt by all molecules.

1. X = Electronegative Atom (O, N, Cl, Br, F, etc.)

Practice Questions



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Transcript

3.6 Predicting Molecular Shape

Learning Objectives

- Utilising VSEPR theory, predict the geometries of molecules through the presence of lone-pairs and electron-dense groups.
- Determine the polarity of a molecule, considering bond types and geometry.

You may have noticed that when drawing or representing molecules, we sometimes draw them with unexpected geometry. Some of the most well-known organic molecules, such as methane, formaldehyde and carbon dioxide all have different shapes.

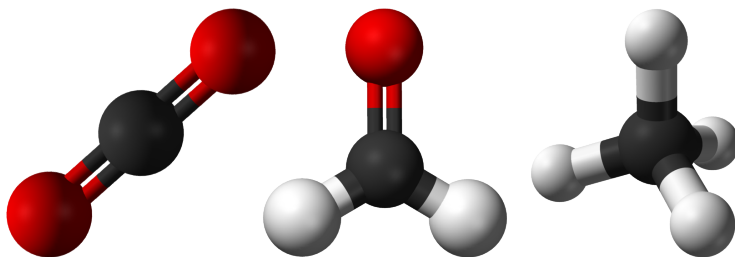


Figure 3.6.1: Carbon dioxide, formaldehyde and methane are all common organic compounds, with vastly different shapes from one another.

There seems to be sound reasoning behind this as all three compounds have different numbers of atoms present. However, consider carbon dioxide compared to a molecule of water; even though both molecules have the same number of atoms, we represent them differently. Water is found to have a bent structure, while carbon dioxide is perfectly linear (see Figure 3.6.2); why is this the case?

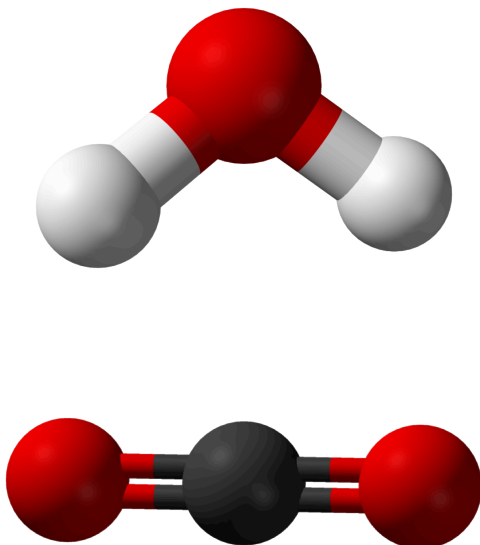


Figure 3.6.2: Water and carbon dioxide have the same number of atoms, but present in different shapes.

VSEPR Theory

To understand why certain molecules present in this way, let us consider the repulsive effect of electrons through the Valence Shell Electron-pair Repulsion model, also known as **VSEPR** theory. Recall that electrons have negative charges. Electrons will, therefore, emit

repelling forces upon each other and will try to roughly position themselves where the minimum repulsive force is experienced. By understanding how many high-density regions of electrons exist, we can best spread them out and predict geometry.

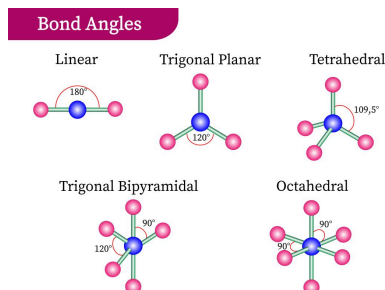


Figure 3.6.3: Different bond angles.
Image attribution: Vector illustration of Bond Angles © aboabdelaah – stock.adobe.com.

In Figure 3.6.3, we observe the preferred orientation of molecules depending on the number of electron-dense regions. The proposed shapes have all electron-dense regions positioned as far as possible from each other. These regions can either be bond pairs or lone electron pairs (on the central atom). In the case of carbon dioxide, we have two regions of

high-electron density, each found at the bonds between the central carbon and oxygen atoms present. As such, a linear formation naturally occurs (see Figure 3.6.4).

In the case of water, however, we detect the two electron-dense groups from the two O-H bonds and the two lone pairs of electrons on the oxygen atoms.

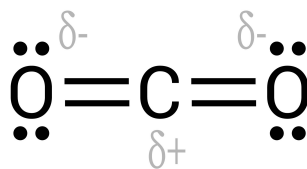


Figure 3.6.4: Lewis dot diagram of carbon dioxide; we see two regions of electron density (at each oxygen).

This means there are four regions of electron density – inferring that the preferred configuration would be a tetrahedral arrangement. Naturally, however, the paired electrons do not generate bonds of their own. Removing two of the electron-dense regions produces a final ‘bent’ or angular geometry (see Figure 3.6.5).

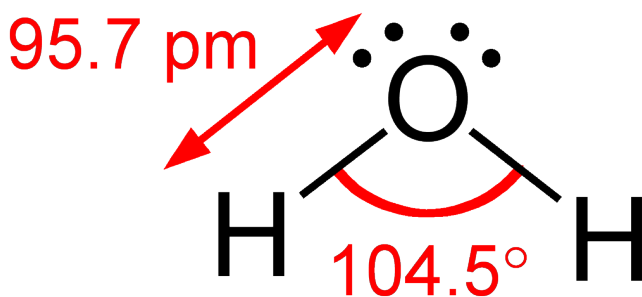


Figure 3.6.5: The bent angle of a water molecule is generated from the lone pairs present.

As such, it is important to note that in VSEPR theory, all molecules take up their preferred orientations – however, the presence of lone pairs makes this appear as if it does not happen. For instance, methane, ammonia and water all have 4 regions of electron density and form tetrahedral arrangements (see Figure 3.6.6). Methane is able to accomplish this, while nitrogen (with one lone pair of electrons) forms a trigonal pyramid without another hydrogen bond. Water (with two lone pairs of electrons) forms a “bent” geometry – there is a clear angle between the two bonds.

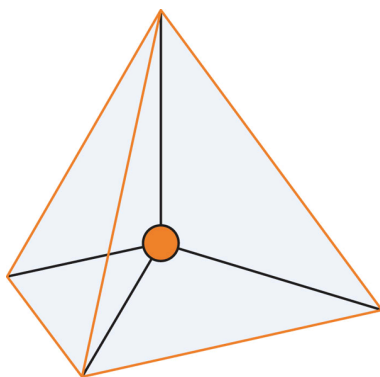


Figure 3.6.6 “Tetrahedral Geometry.” Four electron groups orient themselves in the shape of a tetrahedron. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

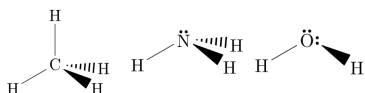


Figure 3.6.7: Methane, ammonia and water. As the number of lone pairs of electrons increases, the shape changes. Transcript.

Therefore, the overall geometry of the molecule can be categorised by the number of atoms and the number of lone electron pairs present on the central atom. Figure 3.6.8 details a helpful guide to

predict shape and can be relied upon as a general rule for most molecules. Let's use the table to help us ascertain the geometry of another organic molecule: formaldehyde.

Number of Electron Groups	Lone Pairs = 0	Lone Pairs = 1	Lone Pairs = 2	Lone Pairs = 3	Lone Pairs = 4
2	 Linear				
3	 Trigonal Planar	 Angular or Bent			
4	 Tetrahedral	 Trigonal Pyramidal	 Angular or Bent		
5	 Trigonal Bipyramidal	 Seesaw	 T-shaped	 Linear	
6	 Octahedral	 Square Pyramidal	 Square Planar	 T-shaped	 Linear

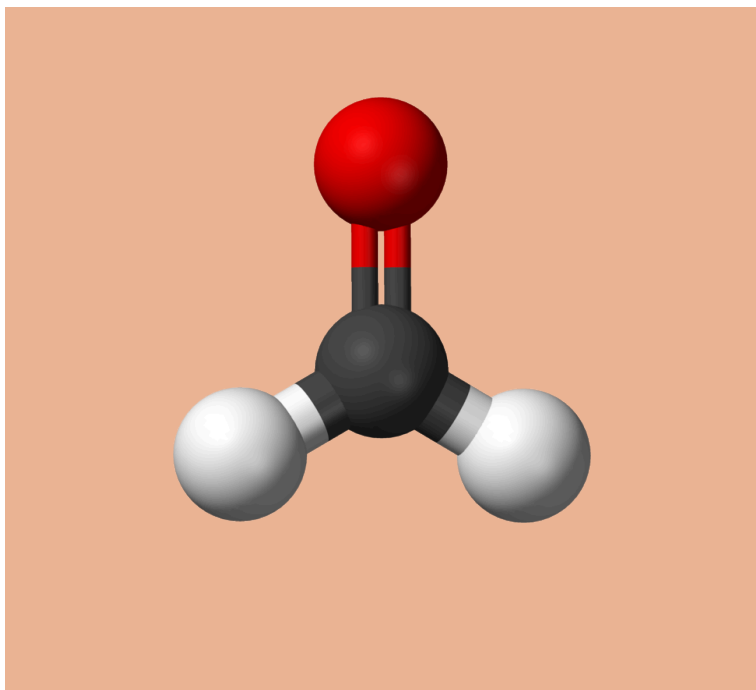
Figure 3.6.8: Table of Molecular Geometry. Image attribution: vector illustration of VSEPR Theory CHART. © aboabdelah – stock.adobe.com. Transcript.

Example

Formaldehyde (CH_2O) is represented in the following Lewis structure:



As such, we observe three electron-dense areas, one for each of the C-H bonds, and one for the C-O bond (note how a double bond is still counted as a single region). There are no lone pairs on the central carbon. Therefore, the preferred geometry is trigonal planar.



VSPER theory represents an idealised set of interactions between electron-dense areas within a molecule. Bond angles and lengths differ slightly between molecules dependant on the types of elements within. However, they are a great generalisation and should be used moving forward in undergraduate studies. Have a look at the simulation below for how real molecules differ to the model predictions.



One or more interactive elements has been excluded from this version of the text. You can view them online here: <https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=1299>

Now that we understand the geometry of molecules, we can begin to determine if a molecule is **polar** or **non-polar**. Recall that these terms were used when discussing bond formation. Polarity refers to the presence of a (whole or partial) magnetic charge generated through atomic interactions, which can be determined through a difference in electronegativity.

Carbon dioxide is a linear molecule consisting of two double-bonded C=O's. These bonds are polar covalent due to their difference in electronegativity. The carbon is electropositive, while the oxygens are electronegative. You would therefore expect the entire molecule to be considered polar. However, if we draw out these negative forces as charges, they directly oppose each other – cancelling each other out, as seen in Figure 3.6.9.

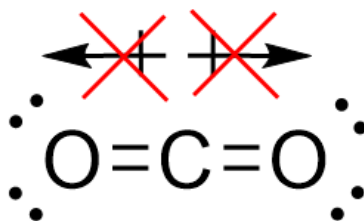
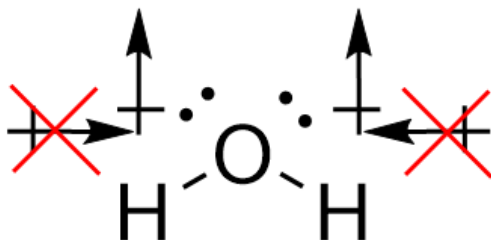
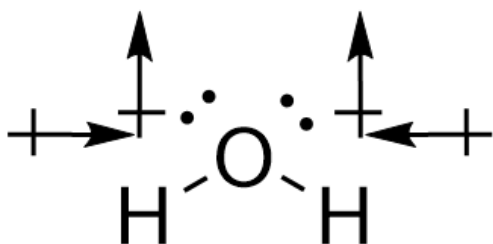
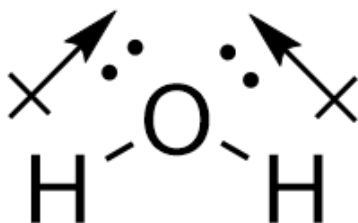


Figure 3.6.9: Although the C=O bonds are polar, they directly oppose each other. The net result is that the molecule remains non-polar. Transcript.

As such, while there are polar covalent bonds with electronegative ends, the entire molecule does not experience a net charge. This makes carbon dioxide non-polar.

Conversely, for water, the O-H bonds are also polar covalent. However, the introduction of the angular bend means that the forces experienced are not directly opposing each other. While the horizontal components of the force do cancel each other out, the

vertical vectors ensure that a permanent dipole is present, meaning that the molecule is polar.



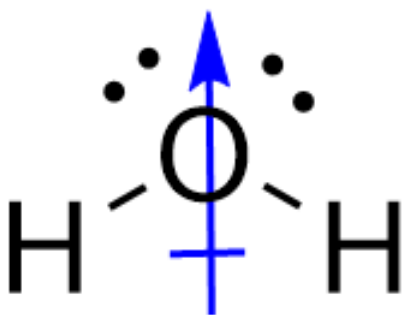


Figure 3.6.10: Determining the polarity/net dipole of a molecule requires us to look at the horizontal and vertical components of the forces at play – cancelling out those that directly oppose. Transcript.

To understand how the polarity of a molecule is determined by the electronegativity of atoms, have a look at the following interactive activity's “Three Atoms” simulation with partial charges enabled:



One or more interactive elements has been excluded from this version of the text. You can view them online here: <https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=1299>

Simulation by PhET Interactive Simulations, University of Colorado

Key Takeaways

- VSEPR Theory allows for the prediction of molecule geometry according to the number of electron-dense regions and lone pair electrons present.
- Molecular geometry can result in certain molecules, which feature polar bonds, to be non-polar due to cancelling out forces.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

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- vector illustration of VSEPR Theory CHART. © aboabdelah - stock.adobe.com is licensed under a All Rights Reserved license

PART IV

CHAPTER 4: CHEMICAL REACTIONS AND EQUATIONS

4.1 Chemical Equations

Learning Objectives

- Define *chemical equation*.
- Identify the parts of a chemical equation.

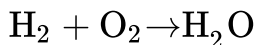
A chemical reaction expresses a chemical change. For example, one chemical property of hydrogen is that it will react with oxygen to make water. This can be expressed as:

Hydrogen reacts with oxygen to make water

We can represent this chemical change more succinctly as:

hydrogen + oxygen \rightarrow water

In this case, the + sign means that the two substances interact chemically with each other, and the \rightarrow symbol implies that a chemical reaction takes place. But substances can also be represented by chemical formulas. Remembering that hydrogen and oxygen, both exist as **diatomic molecules**, we can rewrite our chemical change as:



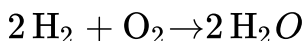
This is an example of a **chemical equation**, which is a concise way of representing a chemical reaction. The initial substances are called **reactants**, and the final substances are called **products**.

Unfortunately, it is also an incomplete chemical equation. **The law of conservation of matter** says that matter cannot be created or destroyed. In chemical equations, the number of atoms of each element in the reactants must be the same as the number of atoms

of each element in the products. If we count the number of hydrogen atoms in the reactants and products, we find two hydrogen atoms. But if we count the number of oxygen atoms in the reactants and products, we find that there are two oxygen atoms in the reactants but only one oxygen atom in the products.

What can we do? Can we change the subscripts in the formula for water so that it has two oxygen atoms in it? No, you *cannot* change the formulas of individual substances because the chemical formula for a given substance is a representation of the chemical composition of that substance.

What you *can* do, however, is to change the number of molecules that react or are produced. We do this one element at a time, going from one side of the reaction to the other, changing the number of molecules of a substance until all elements have the same number of atoms on each side. In our example, the **balanced chemical equation** for the reaction between hydrogen and water is as follows:



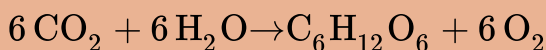
Main Components of a Chemical Equation

- **Reactants:** the starting materials of the reaction. Reactants are written on the left side of the equation.
- **Products:** the materials obtained at the end of the reaction and written on the right side of the equation.
- **An arrow** is drawn between the reactants and products, pointing towards the products to indicate a chemical change.
- **Conditions required for the reaction to occur:** these include pressure and temperature and are written above the arrow.
- **Physical state of the substances involved:** this is written next to the chemical formulas using brackets and abbreviations, as shown next. To illustrate Solid – (s), Liquid – (l), Gas – (g), Aqueous solution (a substance dissolved in water) – (aq).

- **Coefficient:** the numbers written before the chemical formulas to balance the chemical equation.

Examples

Let's take the photosynthesis reaction that occurs in plants as an example. Photosynthesis produces glucose and oxygen in the presence of a pigment known as chlorophyll using sunlight, carbon dioxide and water, as shown in the equation below.



According to the photosynthesis equation shown above:

Reactants: CO_2 , H_2O as they are the starting materials and are written on the left side of the arrow.

Products: $\text{C}_6\text{H}_{12}\text{O}_6$, O_2 as these are the ending products and written on the right side of the arrow.

The conditions required for the reaction to occur are energy and chlorophylls.

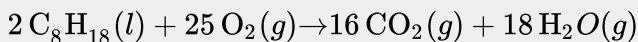
The physical states of the reactants and products are written next to the chemical formula of reactants and products. For instance: CO_2 is gaseous state, and the product, glucose in a solid state.

According to the balanced chemical equation, six CO_2 molecules react with six H_2O molecules to produce one molecule of $\text{C}_6\text{H}_{12}\text{O}_6$ and six molecules of O_2 .

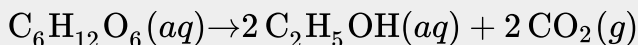
Chemical Reactions in Everyday Life

Chemical reactions are an integral part of our everyday lives, and they occur in various forms and settings, from the functioning of our bodies to the products we use and consume. Here are some examples of chemical reactions along with their corresponding chemical equations:

- **Combustion in a car engine:** when you start your car, gasoline (hydrocarbons) combine with oxygen to produce energy, carbon dioxide, and water vapour, as shown in the following equation.



- **Baking bread:** during the baking of bread, yeast consumes glucose and produces ethanol and carbon dioxide gas, causing the dough to rise, as displayed in the equation below.



- **Rusting of iron:** when iron is exposed to moisture and oxygen, it forms iron oxide (rust):





Figure 4.1.1 Rusty bicycle. Image attribution: old rusty vintage bicycle © sakhorn38-stock.adobe.com.

Key Takeaways

- A chemical equation is a concise description of a chemical reaction.
- A chemical equation consists of several components, such as reactants, products, an arrow, chemical formulas, etc., each serving a specific

purpose to describe the reaction.

- Proper chemical equations are balanced.
- Chemical reactions play a diverse and essential role in our daily lives.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=151#h5p-29>

Transcript

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4.2 Writing and Balancing Chemical Equations

Learning Objectives

- Write chemical equations using chemical formulas.
- Balance chemical equations.

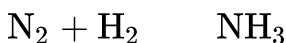
Writing and balancing chemical equations is a fundamental aspect of chemistry, as it allows us to describe and understand chemical reactions in a concise and quantitative manner. Chemical equations consist of reactants, which are the substances that participate in a reaction, and products, which are the substances formed as a result of the reaction. Balancing the equation ensures that the number of atoms of each element is conserved in accordance with the law of conservation of mass.

Here's a step-by-step description of how to write and balance chemical equations:

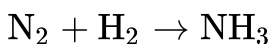
Write the balanced chemical equation for the reaction between nitrogen gas and hydrogen gas to produce ammonia gas.

Step 1: Write the reactants and products of the chemical equation using the correct symbols and formulas.

- Reactants are written on the left side, and products are written on the right side. Use a plus sign to separate each reactant or product from others, as shown below:



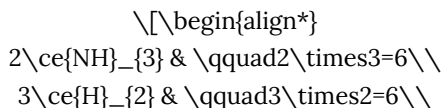
- Place an arrow pointing towards the products between reactants and products, as shown below:

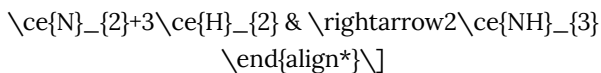


Step 2: Balance the chemical equation using suitable coefficients.

In a balanced chemical equation, an equal number of atoms of each element involved must be present on both sides (i.e. left and right sides of the equation).

- To balance a chemical equation, begin by selecting a specific atom to adjust first. It's often advantageous to commence with the compound with the most significant number of atoms. Within that compound, focus on the element with the highest atom count. In this context, let's initiate the balancing process with the hydrogen found in ammonia. Ammonia consists of three hydrogen atoms, while the left side has only two hydrogen atoms. Six hydrogen atoms are necessary for each side to balance, as this number allows both two and three to be evenly divided. Consequently, add the coefficient "2" in front of ammonia and "3" in front of molecular hydrogen to ensure a total of six hydrogen atoms on both sides of the equation. The process is shown below:





- Next, balance the rest of the atoms in the selected compound. In this example, the next atom will be nitrogen. The right side has two nitrogen atoms, and the left side also has two nitrogen atoms. Therefore, nitrogen is already balanced.
- Balance the remaining elements in the equation, if any. The selected equation only contains two types of elements. Therefore, no more atoms are left to be balanced.

Step 3: Do a final check to ensure all types of atoms in the equation are balanced.

At this point, the equation is balanced for both atoms present. Two nitrogens on each side and six hydrogens on each side.

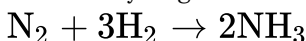


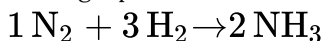
Table 4.2.1 summarises the total number of atoms for each element involved on each side of the equation:

Table 4.2.1 Number of atoms on each side of the equation.

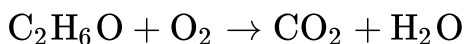
	Left side	Right side
N	2	2
H	$3 \times 2 = 6$	$2 \times 3 = 6$

Step 4: Check whether the coefficients are in their lowest possible whole numbers.

The ratio of **1 : 3 : 2** is the lowest possible in whole numbers, as displayed in the following equation.

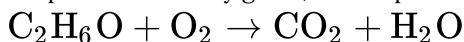


Balance the following chemical equation:



Step 1: Write the reactants and products of the chemical equation using the correct symbols and formulas.

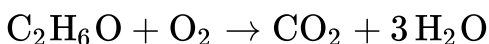
As the chemical equation is already given, this step is not applicable:



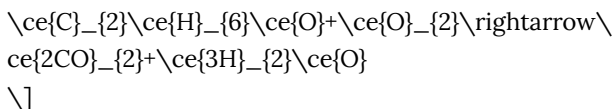
Step 2: Balance the chemical equation using suitable coefficients.

- Choose one atom to balance first. Generally, it helps to start with the compound that is composed of the greatest number of atoms. Focus first on the element of that compound which has the highest number of atoms.

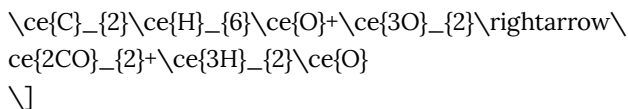
In this case, start with hydrogen. The left side has six hydrogens, and the right side has two hydrogens. Placing three in front of water can give six hydrogen atoms on the right side as well. As can be seen in this equation:



- Next, balance the carbon atoms. The left side contains two carbon atoms, and the right side contains one carbon atom. Using two as the coefficient for carbon dioxide makes two carbons on each side. Thus: \[



- Finally, balance the oxygen atoms. The left side has three oxygen atoms, and the right side has seven oxygen atoms. Placing three in front of molecular oxygen on the left side gives seven oxygen atoms on each side. as can be seen in the following equation:\[



Step 3: Do a final check to ensure all types of atoms in the equation are balanced.

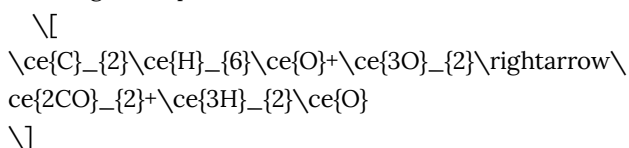
The following Table 4.2.2 summarises the number of atoms of each element present on both sides of the equation.

Table 4.2.2 Number of atoms on each side of the equation.

	Left side	Right side
H	6	$3 \times 2 = 6$
C	2	2
O	$1 + (3 \times 2) = 7$	$(2 \times 2) + 3 = 7$

Step 4: Check whether the coefficients are in their lowest possible whole numbers.

The ratio of **1 : 3 : 2 : 3** is the smallest possible whole number for the given equation. Therefore:



Examples

Write and balance the chemical equation

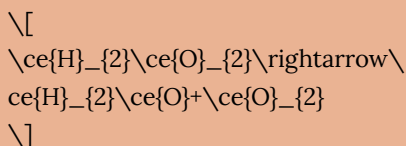
for the decomposition of hydrogen peroxide in the presence of light to water and oxygen gas.

Step 1: Write the reactants and products of the chemical equation using the correct symbols and formulas.

Reactants: Hydrogen peroxide

Products: Water and oxygen gas

Write the reactants on the left side and products on the right side of the equation. Then, place an arrow pointing towards the products as shown by the following equation.



Step 2: Balance the chemical equation using suitable coefficients.

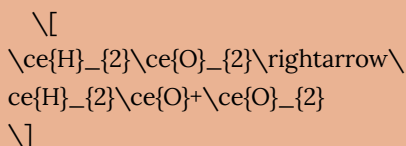
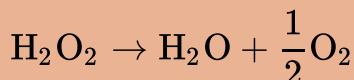


Table 4.2.3 Number of atoms on each side of the equation.

	Left side	Right side
H	2	2
O	2	3

According to the table 4.2.3, hydrogen is already

balanced. There are two oxygen atoms on the left side and three oxygen atoms on the right side. The number of oxygen atoms can be balanced by placing $\frac{1}{2}$ in front of molecular oxygen. This gives two oxygen atoms on each side. Thus:



Step 3: Do a final check to ensure all types of atoms in the equation are balanced.

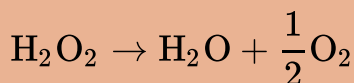
At this point, the equation is balanced for both atoms present: two hydrogens on both sides and two oxygens on both sides, as shown in table 4.2.4.

Table 4.2.4 Number of atoms on each side of the equation.

	Left side	Right side
H	2	2
O	2	2

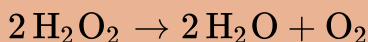
Step 4: Check whether the coefficients are in their lowest possible whole numbers.

The balanced chemical equation obtained in step 3 is as follows:



Except for the coefficient of oxygen on the right side of the equation, the rest have whole-number coefficients. To

make all coefficients in their smallest whole number,
multiply all coefficients by two as has been done here:



Watch the following simulation about balancing chemical equations.



One or more interactive elements has been excluded
from this version of the text. You can view them online

here:

[https://rmit.pressbooks.pub/](https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=902#iframe-phet-1)

[rmitchemistrybridgingcourse/?p=902#iframe-phet-1](https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=902#iframe-phet-1)

Key Takeaways

- When writing chemical equations, place the reactants on the left side of the arrow and the products on the right side of the arrow.
- When balancing chemical equations, it's often helpful to commence with the compound that has the most significant number of atoms.
- You can only change coefficients, not subscripts, in chemical formulas to balance the equation.

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

4.3 Classes of Chemical Reactions: Single and Double Displacement Reactions

Learning Objectives

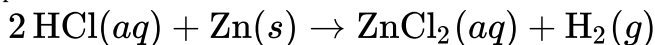
- Recognise chemical reactions as single-replacement reactions and double-replacement reactions.
- Use the periodic table, an activity series, or solubility rules to predict whether single-replacement reactions or double-replacement reactions will occur.

Up to now, we have presented chemical reactions as a topic, but we have not discussed how the products of a chemical reaction can be predicted. Here, we will begin our study of certain types of chemical reactions that allow us to predict what the products of the reaction will be.

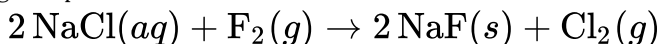
Single-Replacement Reactions

A single-replacement reaction is a chemical reaction in which one element is substituted for another element in a compound,

generating a new element and a new compound as products. For example:

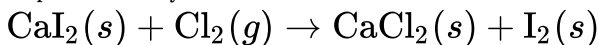


This is an example of a single-replacement reaction. The hydrogen atoms in **HCl** are replaced by **Zn** atoms, and in the process, a new element—hydrogen—is formed. Another example of a single-replacement reaction is:

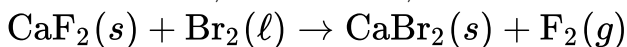


Here, the negatively charged ion changes from chloride to fluoride. A typical characteristic of a single-replacement reaction is that there is one element as a reactant and another element as a product.

Not all proposed single-replacement reactions will occur between two given reactants. This is most easily demonstrated with fluorine, chlorine, bromine, and iodine. Collectively, these elements are called **halogens** and are in the next-to-last column on the periodic table (see Figure 4.3.1 “Halogens on the Periodic Table”). The elements on top of the column will replace the elements below them on the periodic table, but not the other way around. Thus, the reaction represented by:



This reaction will occur, but the reaction,



will not because bromine is below fluorine on the periodic table. This is just one of many ways the periodic table helps us understand chemistry.

[illegible]

Figure 4.3.1 “Halogens on the Periodic Table.” The halogens are the elements in the next-to-last column on the periodic table. Image attribution: “Halogens on the Periodic Table” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

Example 4.3.1

Problems

Will a single-replacement reaction occur? If so, identify the products.

1. $\text{MgCl}_2 + \text{I}_2 \rightarrow ?$
2. $\text{CaBr}_2 + \text{F}_2 \rightarrow ?$

Solutions

1. Because iodine is below chlorine on the periodic table, a single-replacement reaction will not occur.
2. Because fluorine is above bromine on the periodic table, a single-replacement reaction will occur, and the products of the reaction will be CaF_2 and Br_2 .

Test Yourself

Will a single-replacement reaction occur? If so, identify the products.



Answer

Yes; FeCl_2 and I_2

Chemical reactivity trends are easy to predict when replacing anions in simple ionic compounds—simply use their relative positions on the periodic table. However, when replacing the cations, the trends are not as straightforward. This is partly because there are so many elements that can form cations; an element in one column on the periodic table may replace another element nearby, or it may not. A list called the activity series does the same thing the periodic table does for halogens: it lists the elements that will replace other elements below them in single-replacement reactions. A simple activity series is shown below:

Activity Series for Cation Replacement in Single-Replacement Reactions

- Li
- K
- Ba
- Sr
- Ca
- Na
- Mg
- Al
- Mn
- Zn
- Cr
- Fe
- Ni
- Sn
- Pb
- H₂
- Cu
- Hg
- Ag
- Pd
- Pt
- Au

Using an activity series is similar to using the positions of the halogens on the periodic table. **An element on top will replace an element below it in compounds undergoing a single-replacement reaction.** Elements will not replace elements above them in compounds.

Example 4.3.2

Problems

Use the activity series to predict the products, if any, of each equation.

1. $\text{FeCl}_2 + \text{Zn} \rightarrow ?$
2. $\text{HNO}_3 + \text{Au} \rightarrow ?$

Solutions

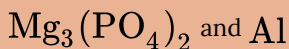
1. Because zinc is above iron in the activity series, it will replace iron in the compound. The products of this single-replacement reaction are ZnCl_2 and Fe .
2. Gold is below hydrogen in the activity series. As such, it will not replace hydrogen in a compound with the nitrate ion. No reaction is predicted.

Test Yourself

Use the activity series to predict the products, if any, of this equation.

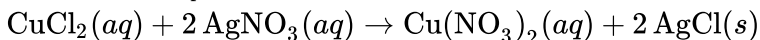


Answer



Double-Replacement Reactions

A double-replacement reaction occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is:



There are two equivalent ways of considering a double-replacement equation: either the cations are swapped, or the anions are swapped. You cannot swap both; you would end up with the same substances you started with. Either perspective should allow you to predict the proper products as long as you pair a cation with an anion and not a cation with a cation or an anion with an anion.

Example 4.3.3

Problem

Predict the products of this double-replacement equation: $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow ?$

Solution

Thinking about the reaction as either switching the cations or switching the anions, we would expect the products to be BaSO_4 and NaCl .

Test Yourself

Predict the products of this double-replacement equation: $\text{KBr} + \text{AgNO}_3 \rightarrow ?$

Answer

KNO_3 and AgBr

Precipitation Reactions and Solubility Rules

Predicting whether a double-replacement reaction will occur is somewhat more difficult than predicting a single-replacement reaction. However, there is one type of double-replacement reaction that we can predict: **the precipitation reaction**. A precipitation reaction occurs when two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve; this new compound falls out of the solution as a solid precipitate. The formation of a solid precipitate is the driving force that makes the reaction proceed.

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble or insoluble). Tables 4.3.1 and 4.3.2 list some general solubility rules. These rules highlight that we need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules in Tables 4.3.1 and 4.3.2. If a compound is soluble, we use the (aq) label with it, indicating that it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of the solution. If all possible products are soluble in a reaction, then no precipitation reaction will be expected.

Table 4.3.1 Some Useful Solubility Rules.

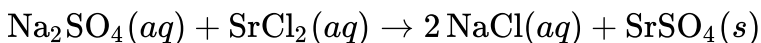
These compounds generally dissolve in water (are soluble):	Exceptions:
All compounds of Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , and NH_4^+	None
All compounds of NO_3^- and $\text{C}_2\text{H}_3\text{O}_2^-$	None
Compounds of Cl^- , Br^- , I^-	Ag^+ , Hg_2^{2+} , Pb^{2+}
Compounds of SO_4^{2-}	Hg_2^{2+} , Pb^{2+} , Sr^{2+} , Ba^{2+}

Table 4.3.2 Some More Useful Solubility Rules.

These compounds generally do not dissolve in water (are insoluble):	Exceptions:
Compounds of CO_3^{2-} and PO_4^{3-}	Compounds of Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , and NH_4^+
Compounds of OH^-	Compounds of Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , NH_4^+ , Sr^{2+} , and Ba^{2+}

To illustrate the rules mentioned in the previous paragraph:

- consider the possible double-replacement reaction between Na_2SO_4 and SrCl_2 . The solubility rules say that all ionic sodium compounds are soluble, and all ionic chloride compounds are soluble except for Ag^+ , Hg^{2+} , and Pb^{2+} , which are not being considered here. Therefore, Na_2SO_4 and SrCl_2 are both soluble.
- The possible double-replacement reaction products are NaCl and SrSO_4 . Are these soluble? NaCl is (by the same rule we just quoted), but what about SrSO_4 ? Compounds of the sulphate ion are generally soluble, but Sr^{2+} is an exception: we expect it to be insoluble—a precipitate. Therefore, we expect a reaction to occur, and the balanced chemical equation would be:



- You would expect to see a visual change corresponding to SrSO_4 precipitating out of solution (Figure 4.3.2 “Double-Replacement Reactions”).



Figure 4.3.2 “Double-Replacement Reactions.” Some double-replacement reactions are obvious because you can see a solid precipitate coming out of the solution. Image attribution: Copper solution © by Choij © Public Domain.

Example 4.3.4

Problems

Will a double-replacement reaction occur? If so, identify the products.

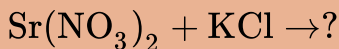
1. $\text{Ca}(\text{NO}_3)_2 + \text{KBr} \rightarrow ?$
2. $\text{NaOH} + \text{FeCl}_2 \rightarrow ?$

Solutions

1. According to the solubility rules, both $\text{Ca}(\text{NO}_3)_2$ and KBr are soluble. Now, we consider what the double-replacement products would be by switching the cations (or the anions)—namely, CaBr_2 and KNO_3 . However, the solubility rules predict that these two substances would also be soluble so that no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and FeCl_2 are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $\text{Fe}(\text{OH})_2$. NaCl is soluble, but according to the solubility rules, $\text{Fe}(\text{OH})_2$ is not. Therefore, a reaction would occur, and $\text{Fe}(\text{OH})_2(s)$ would precipitate out of the solution. The balanced chemical equation is:
$$2\text{NaOH}(aq) + \text{FeCl}_2(aq) \rightarrow 2\text{NaCl}(aq) + \text{Fe}(\text{OH})_2(s)$$

Test Yourself

Will a double-replacement equation occur? If so, identify the products.



Answer

No reaction; all possible products are soluble.

Key Takeaways

- A single-replacement reaction replaces one element for another in a compound.
- The periodic table or an activity series can help predict whether a single-replacement reaction will occur.
- A double-replacement reaction exchanges the cations (or the anions) of two ionic compounds.
- A precipitation reaction is a double-replacement reaction in which one product is a solid precipitate.
- Solubility rules are used to predict whether some double-replacement reactions will occur.

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=155#h5p-33>

Transcript

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4.4 Composition, Decomposition, and Combustion Reactions

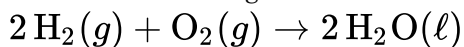
Learning Objectives

- Recognise composition, decomposition, and combustion reactions.
- Predict the products of a combustion reaction.

Three classifications of chemical reactions will be reviewed in this section. Predicting the products in some of them may be difficult, but the reactions are still easier to recognise.

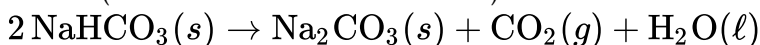
Composition and Decomposition Reactions

A **composition reaction** (sometimes also called a *combination reaction* or a *synthesis reaction*) produces a single substance from multiple reactants. A single substance as a product is the key characteristic of a composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction. Let's consider the following reaction:



In this reaction, water is produced from hydrogen and oxygen. Although there are two molecules of water being produced, there is only one substance—water—as a product. So, this is a composition reaction.

A **decomposition reaction** starts from a single substance and produces more than one substance; that is, it decomposes. One substance as a reactant and more than one substance as a product are the key characteristics of a decomposition reaction. For example, if we look at the decomposition of sodium hydrogen carbonate (also known as sodium bicarbonate):



We can see that sodium carbonate, carbon dioxide, and water are produced from the single substance sodium hydrogen carbonate.

Composition and decomposition reactions are difficult to predict, but recognising them is not so difficult.

Example 4.4.1

Problems

Identify each equation as a composition reaction, a decomposition reaction, or neither.

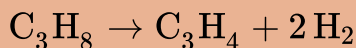
1. $\text{Fe}_2\text{O}_3 + 3 \text{SO}_3 \rightarrow \text{Fe}_2(\text{SO}_4)_3$
2. $\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3$
3. $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 \rightarrow \text{Cr}_2\text{O}_3 + 4 \text{H}_2\text{O} + \text{N}_2$

Solutions

1. In this equation, two substances combine to make a single substance. This is a composition reaction.
2. Two different substances react to make two new substances. This does not fit the definition of either a composition reaction or a decomposition reaction, so it is neither. In fact, you may recognize this as a double-replacement reaction.
3. A single substance reacts to make multiple substances. This is a decomposition reaction.

Test Yourself

Identify the equation as a composition reaction, a decomposition reaction, or neither.



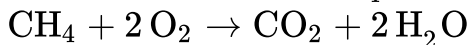
Answer

Decomposition.

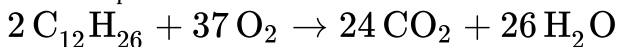
Combustion Reactions

A combustion reaction occurs when a reactant combines with oxygen many times from the atmosphere and produces oxides of all other elements as products; any nitrogen in the reactant is converted to elemental nitrogen, N_2 . Many reactants, called *fuels*, contain mostly carbon and hydrogen atoms, reacting with oxygen

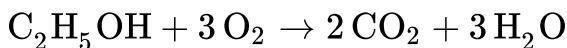
to produce CO_2 and H_2O . For example, the balanced chemical equation for the combustion of methane, CH_4 , is as follows:



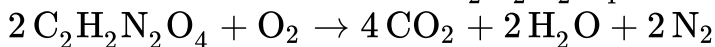
Kerosene can be approximated with the formula $\text{C}_{12}\text{H}_{26}$, and its combustion equation is:



Sometimes, fuels contain oxygen atoms, which must be counted when balancing a chemical equation. One common fuel is ethanol, $\text{C}_2\text{H}_5\text{OH}$, whose combustion equation is:



If nitrogen is present in the original fuel, it is converted to N_2 , not to a nitrogen-oxygen compound. Thus, for the combustion of the fuel dinitroethylene, whose formula is $\text{C}_2\text{H}_2\text{N}_2\text{O}_4$, we have:



Example 4.4.2

Problems

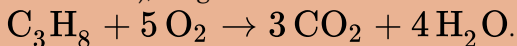
Complete and balance each combustion equation.

1. the combustion of propane, C_3H_8
2. the combustion of ammonia, NH_3

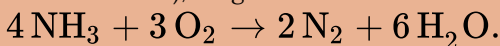
Solutions

1. The products of the reaction are CO_2 and H_2O , so our unbalanced equation is

$\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$. Balancing (and you may have to go back and forth a few times to balance this), we get



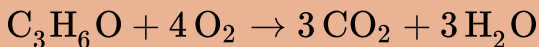
2. The nitrogen atoms in ammonia will react to make N_2 , while the hydrogen atoms will react with O_2 to make H_2O , thus $\text{NH}_3 + \text{O}_2 \rightarrow \text{N}_2 + \text{H}_2\text{O}$. To balance this equation without fractions (which is the convention), we get



Test Yourself

Complete and balance the combustion equation for cyclopropanol, $\text{C}_3\text{H}_6\text{O}$.

Answer





Propane is a fuel used to provide heat for some homes. Propane is stored in large tanks like the one shown here. Image attribution: Propane Tank © vistavision © CC BY-NC-ND (Attribution NonCommercial NoDerivs).

Key Takeaways

- A composition reaction produces a single substance from multiple reactants.

- A decomposition reaction produces multiple products from a single reactant.
- Combustion reactions occur when substances react with oxygen gas.

Exercises

Practice Questions



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<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=162#h5p-34>

Transcript

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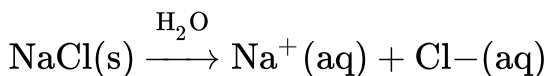
4.5 Ionic Equations

Learning Objectives

- Write ionic equations for chemical reactions between ionic compounds.
- Recognise spectator ions.
- Write net ionic equations for chemical reactions between ionic compounds.

As we learned in Chapter 3, ionic compounds are compounds between metals and nonmetals or compounds that contain recognisable polyatomic ions. Now, we take a closer look at reactions that include ionic compounds.

An important distinction between ionic compounds and molecular compounds is how they dissolve in a liquid such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process—for example, with **NaCl**:



When **NaCl** dissolves in water, the ions separate and go their own way in the solution; the ions are now written with their

respective charges, and the (aq) phase label emphasises that they have been dissolved (Figure 4.3.1 “Ionic Solutions”). This process is called dissociation; we say that the ions *dissociate*.

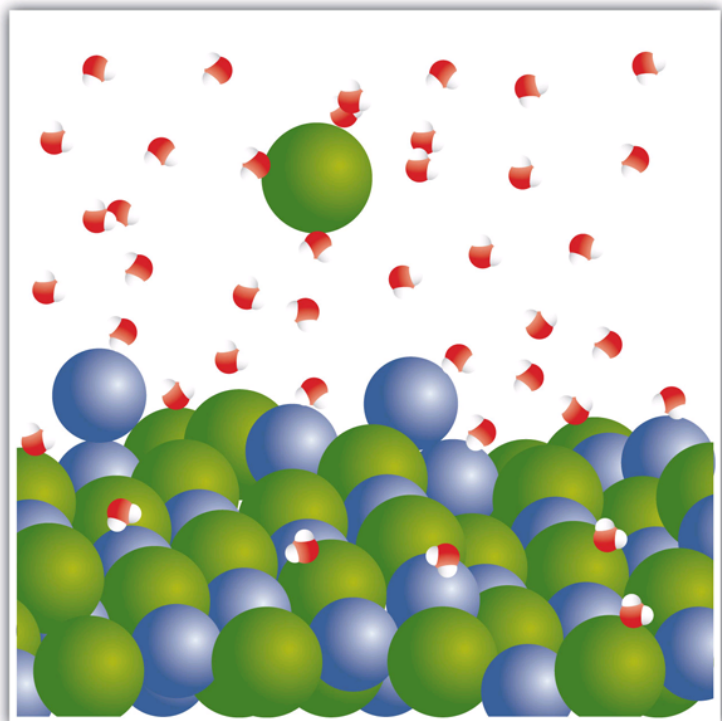
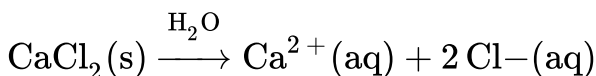
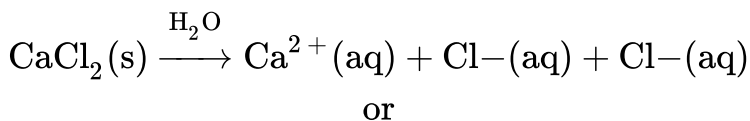


Figure 4.5.1 “Ionic Solutions.” When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution. Image attribution: “Ionic Solutions” by David W. Ball © CC BY-NC-SA (Attribution NonCommercial ShareAlike).

The behaviour of ionic compounds was first suggested

by the Swedish chemist Svante August Arrhenius (1859–1927) as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry.

All ionic compounds that dissolve behave this way. Keep in mind that when the ions separate, *all* the ions separate. Thus, when CaCl_2 dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from each other:



That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions); they do not stick together to make Cl_2^- or Cl_2^{2-} . They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

Example 4.5.1

Problems

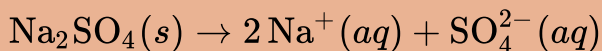
Write the chemical equation that represents the dissociation of each ionic compound.

1. KBr
2. Na_2SO_4

Solution

1. $\text{KBr}(\text{s}) \rightarrow \text{K}^+(\text{aq}) + \text{Br}^-(\text{aq})$
2. Sodium ions go their own way, but the sulphate ion

stays together as the sulphate ion. The dissolving equation is:



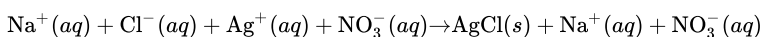
Test Yourself

Write the chemical equation that represents the dissociation of $(\text{NH}_4)_2\text{S}$.

Answer



When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when $\text{NaCl}(aq)$ reacts with $\text{AgNO}_3(aq)$ in a double-replacement reaction to precipitate $\text{AgCl}(s)$ and form $\text{NaNO}_3(aq)$, the complete ionic equation includes NaCl , AgNO_3 , and NaNO_3 written as separated ions:



This is more representative of what is occurring in the solution.

Example 4.5.2

Problems

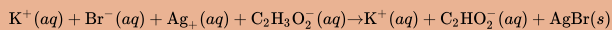
Write the complete ionic equation for each chemical reaction.

1. $\text{KBr}(aq) + \text{AgC}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{KC}_2\text{H}_3\text{O}_2(aq) + \text{AgBr}(s)$
2. $\text{MgSO}_4(aq) + \text{Ba}(\text{NO}_3)_2(aq) \rightarrow \text{Mg}(\text{NO}_3)_2(aq) + \text{BaSO}_4(s)$

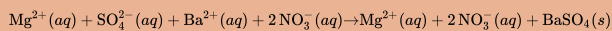
Solution

For any ionic compound that is aqueous, we write the compound as separated ions.

1. The complete ionic equation is:

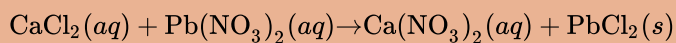


2. The complete ionic equation is:

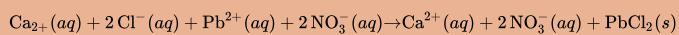


Test Yourself

Write the complete ionic equation for:

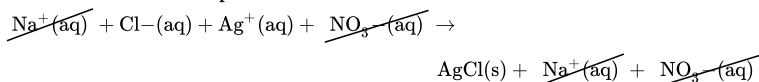


Answer

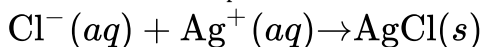


You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in:

$\text{Na}^+(aq) + \text{Cl}^-(aq) + \text{Ag}^+(aq) + \text{NO}_3^-(aq) \rightarrow \text{AgCl}(s) + \text{Na}^+(aq) + \text{NO}_3^-(aq)$
 the $\text{Ag}^+(aq)$ and $\text{Cl}^-(aq)$ ions become $\text{AgCl}(s)$, but the $\text{Na}^+(aq)$ ions and the $\text{NO}_3^-(aq)$ ions stay as $\text{Na}^+(aq)$ ions and $\text{NO}_3^-(aq)$ ions. These two ions are examples of spectator ions, which are ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:



What remains when the spectator ions are removed is called the *net ionic equation*, which represents the actual chemical change occurring between the ionic compounds:



It is important to reiterate that the spectator ions are still present in the solution, but they don't experience any net chemical change, so they are not written in a net ionic equation.

Example 4.5.3

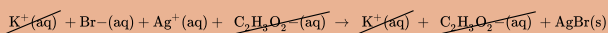
Problems

Write the net ionic equation for each chemical reaction.

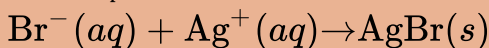
1. $\text{K}^+(aq) + \text{Br}^-(aq) + \text{Ag}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) \rightarrow \text{K}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq) + \text{AgBr}(s)$
2. $\text{Mg}^{2+}(aq) + \text{SO}_4^{2-}(aq) + \text{Ba}^{2+}(aq) + 2\text{NO}_3^-(aq) \rightarrow \text{Mg}^{2+}(aq) + 2\text{NO}_3^-(aq) + \text{BaSO}_4(s)$

Solution

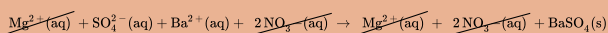
1. In the first equation, the $\text{K}^+(aq)$ and $\text{C}_2\text{H}_3\text{O}_2^-(aq)$ ions are spectator ions, so they are cancelled:



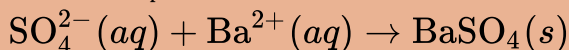
The net ionic equation is:



2. In the second equation, the $\text{Mg}^{2+}(aq)$ and $\text{NO}_3^-(aq)$ ions are spectator ions, so they are cancelled:

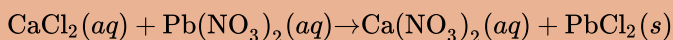


The net ionic equation is:

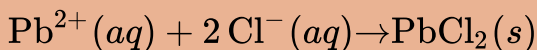


Test Yourself

Write the net ionic equation for:



Answer



Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of the compound that can be dissolved in a sample of water. For example, you can dissolve a maximum of **36.0g** of **NaCl** in **100g** of water at room temperature, but you can dissolve only **0.00019g** of **AgCl** in **100g** of water. We consider **NaCl** soluble but **AgCl** insoluble.

One place where solubility is important is in the tank-type water heater found in many homes. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate (**CaCO₃**). However, **CaCO₃** has the relatively unusual property of being less soluble in hot water than

in cold water. So, as the water heater operates by heating water, CaCO_3 can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.

Key Takeaways

- Ionic compounds that dissolve separately into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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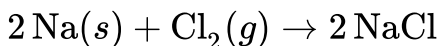
Transcript

4.6 Oxidation-Reduction Reactions

Learning Objectives

- Define oxidation and reduction.
- Identify redox reactions.
- Assign oxidation numbers to atoms in simple compounds.
- Recognise the oxidising agent and the reducing agent in a redox reaction.
- Balancing a redox reaction.

Consider this chemical reaction:



The reactants are elements, and it is assumed that they are electrically neutral; they have the same number of electrons as protons. The product, however, is ionic; it is composed of Na^+ and Cl^- ions. Somehow, the individual sodium atoms as reactants had to lose an electron to make the Na^+ ion, while the chlorine atoms as reactants had to each gain an electron to become the Cl^- ion. This reaction involves the *transfer of electrons* between atoms.

In reality, electrons are lost by some atoms and gained by other atoms simultaneously. However, mentally, we can separate the two processes. Oxidation is defined as the loss of one or more electrons by an atom (Figure 4.6.1). Reduction is defined as the gain of one

or more electrons by an atom (Figure 4.6.1). In reality, oxidation and reduction always occur together; it is only mentally that we can separate them. Chemical reactions that involve the transfer of electrons are called oxidation-reduction (or redox) reactions.

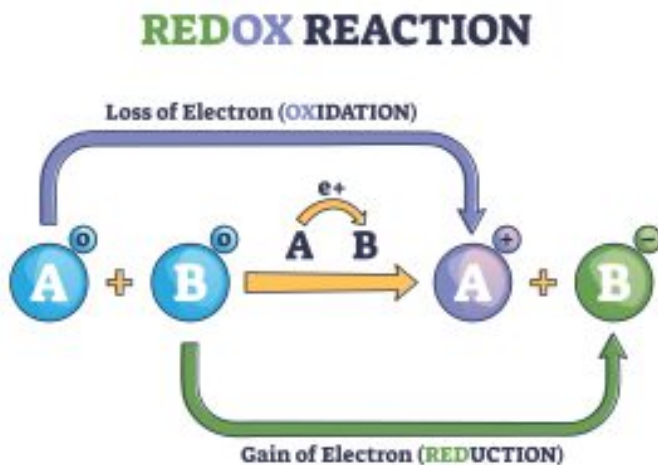


Figure 4.6.1 A schematic representation of Redox reactions. Image attribution: Redox reaction as atoms chemical oxidation states change outline diagram. Labeled educational explanation scheme with electron gain and loss in oxidation or reduction process vector illustration. © VectorMine-stock.adobe.com.

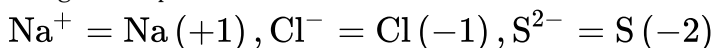
Oxidation Number

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use an artificial count called the oxidation number to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on a series of rules. Oxidation numbers are not necessarily

equal to the charge on the atom; we must keep the concepts of charge and oxidation numbers separate.

The rules for assigning oxidation numbers to atoms are as follows:

1. The oxidation number of an atom as an element is zero.
Examples: O_2 , P_4 , Fe , Br_2 are the elemental forms at room temperature, and they all have an oxidation number of zero.
2. The oxidation number of oxygen is usually -2 (except when bonded to F). Remember that in the element, oxygen (O_2) has an oxidation number of zero, whilst in peroxides (O_2^{2-}) each oxygen atom has an oxidation number of -1 .
3. The oxidation number of a monoatomic ion is equal to the ion's charge. Examples:



4. The algebraic sum of the oxidation numbers in a neutral compound is zero; in a **polyatomic ion**, the sum of the oxidation number is equal to the ion's overall charge.
5. In combination of elements, the more electronegative element has its characteristic negative oxidation number (for example, (-3) for nitrogen, (-2) for oxygen, (-1) for chlorine), and the more electropositive element has a positive oxidation number (for example, $(+1)$ for hydrogen, $(+2)$ for calcium, $(+3)$ for aluminium).
6. The oxidation number of hydrogen is usually $+1$ due to its electronegativity being lower than that of the main-group elements to which it is normally bonded. With more electropositive elements, hydrogen has an oxidation number of (-1) for instance, $\text{NaH} = \text{Na}^+ \text{H}^-$.

Problems

Assign oxidation numbers to the atoms in each substance:

1. Br_2
2. SiO_2
3. $\text{Ba}(\text{NO}_3)_2$

Solutions

1. Br_2 is the elemental form of bromine. Therefore, by rule 1, each atom has an oxidation number of 0.
2. By rule 2, oxygen is normally assigned an oxidation number of -2. For the sum of the oxidation numbers to equal the charge on the species (which is zero), the silicon atom is assigned an oxidation number of +4.
3. The compound barium nitrate can be separated into two parts: the Ba^{2+} ion and the nitrate ion. Considering these separately, the Ba^{2+} ion has an oxidation number of +2 by rule 3. Now consider the NO_3^- ion. Oxygen is assigned an oxidation number of -2, and there are three oxygens. According to rule 4, the sum of the oxidation number on all atoms must equal the charge on the species, so we have the

simple algebraic equation $x + 3(-2) = -1$, where x is the oxidation number of the nitrogen atom and -1 represents the charge on the species. Evaluating,

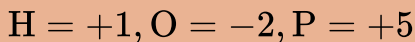
$$\begin{array}{rcl} x + (-6) & = & -1 \\ x & = & 5 \end{array}$$

Thus, the oxidation number on the N atom in the nitrate ion is +5.

Test Yourself

Assign oxidation numbers to the atoms in H_3PO_4 .

Answer



All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When the oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidised*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction are thus also defined in terms of increasing or decreasing oxidation numbers, respectively.

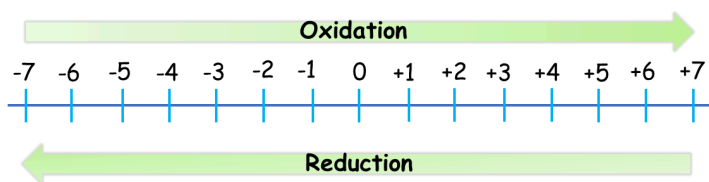
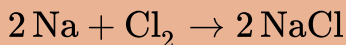


Figure 4.6.2 Change of oxidation number with oxidation and reduction. Oxidation leads to a loss of electrons and an increase in oxidation number. Reduction is a gain of electrons, and so leads to a decrease in oxidation number.

Example 4.6.2

Problem

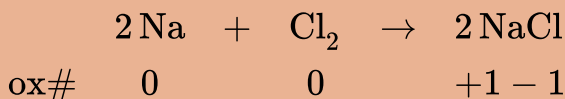
Identify what is being oxidised and reduced in this redox equation:



Solution

Consider the reactants. Because both reactants are the elemental forms of their atoms, the **Na** and **Cl** atoms as

reactants have oxidation numbers of 0. In the ionic product, the sodium ions have an oxidation number of $+1$, while the chloride ions have an oxidation number of -1 , as shown below:



We note that sodium is increasing its oxidation number from 0 to $+1$, so it is being oxidised; chlorine is decreasing its oxidation number from 0 to -1 , so it is being reduced (Figure 4.6.3).

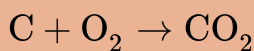


Figure 4.6.3: Electron transfer between reactants and products.

Because oxidation numbers are changing, this is a redox reaction. Note that the total number of electrons lost by the sodium (two, one lost from each atom) is gained by the chlorine atoms (two, one gained for each atom).

Test Yourself

Identify what is being oxidised and reduced in this redox equation:



Answer

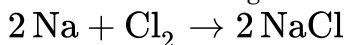
C is being oxidised from 0 to +4; O is being reduced from 0 to -2.

Oxidising Agents and Reducing Agents

Oxidising agents can oxidise another reactant by accepting electrons. Therefore, the oxidising agent becomes reduced (by gaining electrons) while oxidising another reactant (causing it to lose electrons).

Reducing agents reduce another reactant by offering electrons to the other reactant. Thus, the reducing agent becomes oxidised (by losing electrons) while reducing (gaining electrons) another reactant (causing it to gain electrons).

For instance, let's consider the following reaction:



Here, sodium oxidised to become Na^+ by donating an electron to chlorine, which causes chlorine to be reduced. Thus, sodium is the reducing agent in this reaction. Chlorine becomes chloride anion (Cl^-) by accepting an electron from sodium, which makes sodium to be oxidised. Therefore, chlorine is the oxidising agent in this reaction.

Balancing of Redox Equations

Redox equations are usually written as balanced net ionic equations,

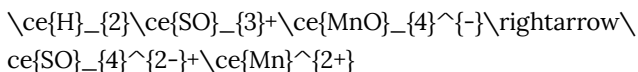
with spectator ions (ions that do not take part in the overall redox process) being omitted.

Determining the **stoichiometry** of a redox reaction can be challenging, even when the products of the reaction are known. For this reason, it is much more convenient and methodical to concentrate on the stoichiometry of each redox “half-reaction” or “redox couple” individually.

A five-step method for balancing redox equations, called the “half-reaction” method, is used, and after mastering these operations, any chemical redox equation can be balanced by combining these five steps.

1. Divide the equation into the appropriate “half-reactions”, which can be balanced separately, for example:

Overall reaction \[

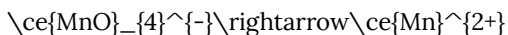


\] Assign an oxidation number to each element in the ions or molecules of reactants and products:

Oxidation number of **Mn** in MnO_4^- is **+7** and **Mn** oxidation number in Mn^{2+} is **+2**

Thus, Mn^{7+} is being reduced to Mn^{2+} .

So, the Reduction half-reaction is: \[



\] Oxidation number of **S** in H_2SO_3 is **+4**, and oxidation number of **S** in SO_4^{2-} ion is **+6**

So, **S** is being oxidised from **+4** \rightarrow **+6**.

So, the oxidation half-reaction is: \[



\]

2. Write a separate balanced chemical equation for each redox-active species. This is the mass balance step: Add water to balance oxygen:

$$\text{MnO}_4^- + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$$
 Add H^+ to balance hydrogens:

$$\text{H}_2\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 4\text{H}^+$$

3. Although we have now achieved mass balance, we also need to balance the charge by adding electrons (negative charge) as needed. This generates the balanced electrochemical half-reaction. The total charge on each side of a balanced reaction must be the same:

$$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$$

Left-hand side (LHS) total charge $+2$

Right-hand side (RHS) total charge $+2$

And:

$$\text{H}_2\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^-$$

LHS total charge 0

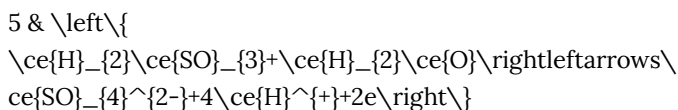
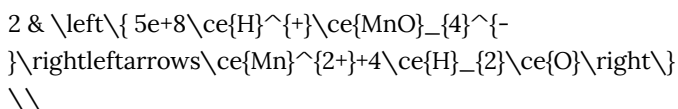
RHS total charge 0

If all three steps have been done correctly, one half-reaction will have electrons on the left-hand side of the equation (the reduction half-reaction), while the other half-reaction will have electrons on the right-hand side of the equation (the oxidation half-reaction).

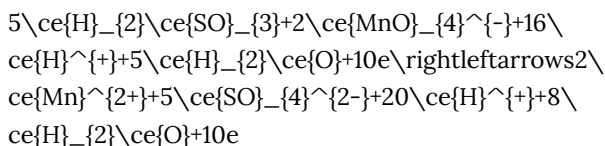
4. The balanced half-reactions can be termed 'electrochemical equations' since they involve the electrons. To obtain a chemical equation for the reaction, we need to eliminate the electrons from the overall equation. Hence, the fourth step is the addition of the two half-reactions in such a way that the electrons cancel out. This may require multiplying one equation or both by an appropriate integer.

For the two half-reactions above:

$$\begin{array}{l} \text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \\ \text{H}_2\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \end{array}$$



The first reaction is multiplied by two, and the second reaction is multiplied by five, and both are added to give:



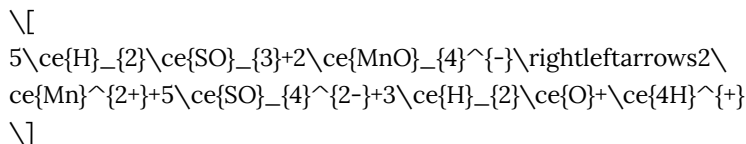
Note the equal number of electrons on each side of the equation cancel out.

- The final step is to cancel the electrons and simplify the chemical equation by eliminating any chemical species common to each side of the equation.

As shown in Table 4.6.1, balance the H^+ ions and H_2O molecules on each side.

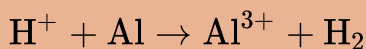
Table 4.6.1: Number of atoms of elements on each side of the equation.

LHS	RHS	Balance
16H^+	20H^+	$20 - 16 = 4\text{H}^+$
$5\text{H}_2\text{O}$	$8\text{H}_2\text{O}$	$8 - 5 = 3\text{H}_2\text{O}$



Test Yourself

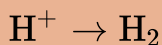
Balance the following redox equation:



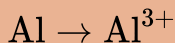
Answer

Step 1 – Divide the equation into the appropriate “half-reactions” which can be balanced separately.

Reduction half-reaction:

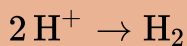


Oxidation half-reaction:

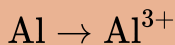


Step 2 – Write a separate balanced chemical equation for each redox-active species. This is the mass balance step.

Reduction half-reaction:

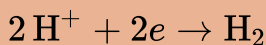


Oxidation half-reaction:



Step 3 – Balance the net charge of each half-reaction by adding electrons.

Reduction half-reaction:

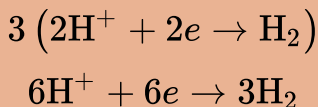


Oxidation half-reaction:

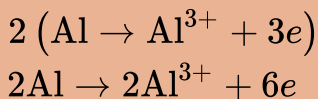


Step 4 – Add two half-reactions in such a way that the electrons cancel out. Multiply the reduction half-reaction by three and the oxidation half-reaction by two to get an equal number of electrons on both sides.

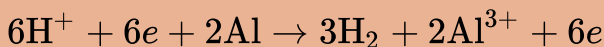
Reduction half-reaction:



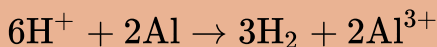
Oxidation half-reaction:



Add two reactions:



Step 5 – Cancel the electrons and simplify the chemical equation by eliminating any chemical species common to each side of the equation as follows:



Batteries

Redox reactions have a wide range of applications across science, industry, and technology. From powering our electronic devices to enabling critical biological processes and environmental cleanup, these reactions are integral to our everyday lives and the advancement of various fields. One of the most well-known applications of redox reactions is in batteries (Figure 4.6.4).



Figure 4.6.4 Different types of batteries. Image attribution: Set of metal batteries, accumulators, button cells, power bank of different sizes. Nickel and lithium energy sources for electric devices. Colored flat vector illustration isolated on white background © Good Studio-stock.adobe.com.

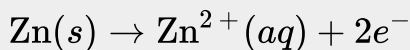
Redox reactions play a critical role in the functioning of batteries as they consist of two electrodes, an anode and a cathode, separated by an electrolyte. When the battery is discharging, a chemical reaction occurs at the anode (oxidation) and another at the cathode (reduction). Electrons flow from the anode to the cathode through an external circuit, creating an electric current. This flow of electrons is what powers our devices and vehicles. The overall reaction in a battery is a redox reaction, and the specific reactions depend on the type of battery (e.g. alkaline, lithium-ion, lead-acid).

For instance, the primary redox reaction that occurs in an alkaline battery involves the following key components:

- *Anode (Negative Electrode)*

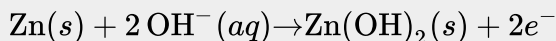
The anode typically consists of powdered zinc (**Zn**) and a gel-like mixture of zinc and potassium hydroxide (**KOH**). The zinc metal serves as the active material in the anode.

The redox reaction at the anode is the oxidation of zinc:



As zinc atoms lose electrons, they are oxidised into zinc ions (Zn^{2+}), releasing two electrons in the process. Zn^{2+} then reacts with hydroxyl ions in the electrolyte solution to form zinc hydroxide.

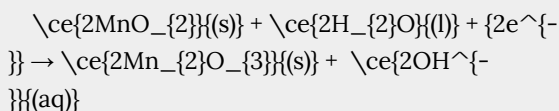
The overall reaction at the anode is:



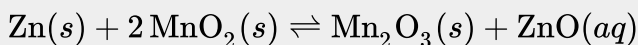
- Cathode (Positive Electrode)

The cathode is usually composed of manganese dioxide (MnO_2) and a mixture of graphite.

The redox reaction at the cathode is the reduction of manganese dioxide:



The overall redox reaction in an alkaline battery is:



Key Takeaways

- Oxidation-reduction (redox) reactions involve the transfer of electrons from one atom to another.
- Oxidation numbers are used to keep track of electrons in atoms.
- There are rules for assigning oxidation numbers to atoms.

- Oxidation is an increase in oxidation number (a loss of electrons); reduction is a decrease in oxidation number (a gain of electrons).
- An oxidising agent causes the oxidation of another reactant by accepting electrons, and a reducing agent causes the reduction of another reactant by donating electrons.
- Redox equations can be balanced using the half-reaction method.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=164#h5p-36>

Transcript

Media Attributions

- Redox reaction as atoms chemical oxidation states change outline diagram. Labeled educational explanation scheme with electron gain and loss in oxidation or reduction process vector illustration. © VectorMine-stock.adobe.com
- Set of metal batteries, accumulators, button cells, power bank of different sizes. Nickel and lithium energy sources for electric devices. Colored flat vector illustration isolated on white background © Good Studio-stock.adobe.com

4.7 Neutralisation Reactions

Learning Objectives

- Identify an acid and a base.
- Identify a neutralisation reaction and predict its products.

Acids and Bases

An acid is any compound that increases the amount of hydrogen ion (H^+) in an aqueous solution. The chemical opposite of an acid is a base. The equivalent definition of a base is that a base is a compound that increases the amount of hydroxide ion (OH^-) in an aqueous solution. These original definitions were proposed by Arrhenius (the same person who proposed ion dissociation) in 1884, so they are referred to as the Arrhenius definitions of an acid and a base, respectively. **These definitions simply state that a compound that generates hydrogen ions when dissolved in water is an acid, and a compound that generates hydroxyl ions when dissolved in water is a base.**

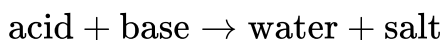
You may recognise that based on the description of a hydrogen atom, an H^+ ion is a hydrogen atom that has lost its lone electron; that is, H^+ is simply a proton. Do we really have bare protons moving about in aqueous solution? No. What is more likely is that the H^+ ion has attached itself to one (or more) water molecule(s).

To represent this chemically, we define the hydronium ion $\text{H}_3\text{O}^+(\text{aq})$, a water molecule with an extra hydrogen ion attached to it, as H_3O^+ , which represents an additional proton attached to a water molecule. We use the hydronium ion as the more logical way a hydrogen ion appears in an aqueous solution, although in many chemical reactions, H^+ and H_3O^+ are treated equivalently.

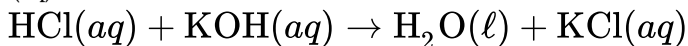
You will learn about acids and bases in detail in chapter 6.

Neutralisation Reactions

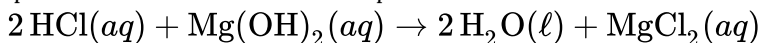
The reaction of an acid and a base is called a neutralisation reaction. Although acids and bases have their own unique chemistries, the acid and base cancel each other's chemistry to produce a rather innocuous substance—water. In fact, the general reaction between an acid and a base is



where the term salt is generally used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base. It should be noted that, in chemistry, the word salt refers to more than just table salt. For example, the balanced chemical equation for the reaction between $\text{HCl}(\text{aq})$ and $\text{KOH}(\text{aq})$ is



where the salt is KCl . By counting the number of atoms of each element, we find that only one water molecule is formed as a product. However, in the reaction between $\text{HCl}(\text{aq})$ and $\text{Mg}(\text{OH})_2(\text{aq})$, additional molecules of HCl and H_2O are required to balance the chemical equation:



Here, the salt is MgCl_2 . This is one of several reactions that

take place when a type of antacid—a base—is used to treat stomach acid.

Example 4.7.1

Problems

Write the neutralisation reactions between each acid and base.

1. $\text{HNO}_3(aq)$ and $\text{Ba}(\text{OH})_2(aq)$
2. $\text{H}_3\text{PO}_4(aq)$ and $\text{Ca}(\text{OH})_2(aq)$

Solutions

First, we will write the chemical equation with the formulas of the reactants and the expected products; then we will balance the equation.

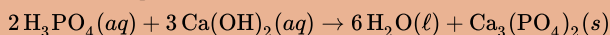
1. The expected products are water and barium nitrate, so the initial chemical reaction is
$$\text{HNO}_3(aq) + \text{Ba}(\text{OH})_2(aq) \rightarrow \text{H}_2\text{O}(\ell) + \text{Ba}(\text{NO}_3)_2(aq)$$

. To balance the equation, we need to realise that there will be two H_2O molecules, so two HNO_3 molecules are required:

$$2\text{HNO}_3(aq) + \text{Ba}(\text{OH})_2(aq) \rightarrow 2\text{H}_2\text{O}(\ell) + \text{Ba}(\text{NO}_3)_2(aq)$$

This chemical equation is now balanced.
2. The expected products are water and calcium

phosphate, so the initial chemical equation is $\text{H}_3\text{PO}_4(aq) + \text{Ca}(\text{OH})_2(aq) \rightarrow \text{H}_2\text{O}(\ell) + \text{Ca}_3(\text{PO}_4)_2(s)$. According to the solubility rules, $\text{Ca}_3(\text{PO}_4)_2$ is insoluble, so it has an (*s*) phase label. To balance this equation, we need two phosphate ions and three calcium ions; we end up with six water molecules to balance the equation:

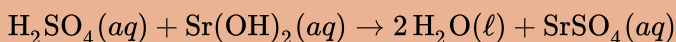


This chemical equation is now balanced.

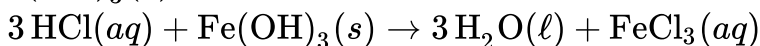
Test Yourself

Write the neutralisation reaction between $\text{H}_2\text{SO}_4(aq)$ and $\text{Sr}(\text{OH})_2(aq)$.

Answer



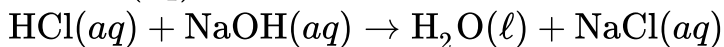
Neutralisation reactions are one type of chemical reaction that proceeds even if one reactant is not in the aqueous phase. For example, the chemical reaction between $\text{HCl}(aq)$ and $\text{Fe}(\text{OH})_3(s)$ still proceeds according to the equation:



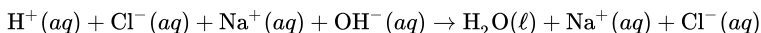
even though $\text{Fe}(\text{OH})_3$ is not soluble. When one realises that $\text{Fe}(\text{OH})_3(s)$ is a component of rust, this explains why some cleaning solutions for rust stains contain acids—the neutralisation reaction produces products that are soluble and wash away.

Complete and net ionic reactions for neutralisation reactions will

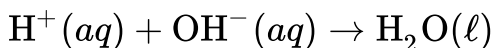
depend on whether the reactants and products are soluble, even if the acid and base react. For example, in the reaction of $\text{HCl}(aq)$ and $\text{NaOH}(aq)$:



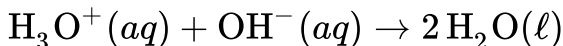
The complete ionic reaction is:



The $\text{Na}^+(aq)$ and $\text{Cl}^-(aq)$ ions are spectator ions, so we can remove them to have:

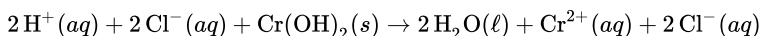


as the net ionic equation. If we wanted to write this in terms of the hydronium ion, $\text{H}_3\text{O}^+(aq)$, we would write it as:

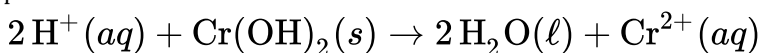


With the exception of the introduction of an extra water molecule, these two net ionic equations are equivalent.

However, for the reaction between $\text{HCl}(aq)$ and $\text{Cr}(\text{OH})_2(s)$, because chromium(II) hydroxide is insoluble, we cannot separate it into ions for the complete ionic equation:



The chloride ions are the only spectator ions here, so the net ionic equation is:



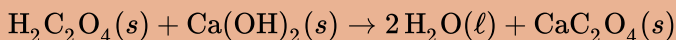
Example 4.7.2

Problem

Oxalic acid, $\text{H}_2\text{C}_2\text{O}_4(s)$, and $\text{Ca}(\text{OH})_2(s)$ react very slowly. What is the net ionic equation between these two substances if the salt formed is insoluble? (The anion in oxalic acid is the oxalate ion, $\text{C}_2\text{O}_4^{2-}$.)

Solution

The products of the neutralisation reaction will be water and calcium oxalate:

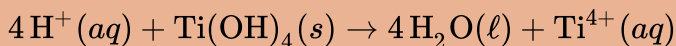


Because nothing is dissolved, there are no substances to separate into ions, so the net ionic equation is the equation of the three solids and one liquid.

Test Yourself

What is the net ionic equation between $\text{HNO}_3(aq)$ and $\text{Ti}(\text{OH})_4(s)$?

Answer



Key Takeaways

- The Arrhenius definition of an acid is a substance that increases the amount of H^+ in an aqueous solution.
- The Arrhenius definition of a base is a substance that increases the amount of OH^- in an aqueous solution.
- Neutralisation is the reaction of an acid and a base, which forms water and a salt.
- Net ionic equations for neutralisation reactions may include solid acids, solid bases, solid salts, and water.

Exercises

Practice Questions





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Transcript

PART V

CHAPTER 5: MEASURING CHEMISTRY

5.1 The Mole and Avogadro's Number

Learning Objectives

- Apply Avogadro's number (6.022×10^{23}) to represent large amounts of a species through the mole within chemical contexts.

In speech, we use alternative units to describe amounts of items. For instance, 1 dozen = 12 objects, 1 score = 20 objects, 1 ream = 500 sheets paper, likewise in chemistry, chemist consider:

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ objects}$$

There are many interesting units and groupings in science! Leigh Mercer is most often credited with a famous mathematical limerick to assist in remembering rather odd values.

$$\frac{12 + 144 + 20 + 3 * \sqrt{4}}{7} + (5 \times 11) = 9^2 + 0$$

A dozen, a gross, and a score

*Plus three times the square root of four
Divided by seven
Plus five times eleven
Is nine squared and not a bit more.*

The objects can be anything, such as houses, sand, eggs, oranges, etc. However, in chemistry, the objects we are concerned with are mainly atoms, ions or molecules – objects that are invisible to the naked eye!

One **mole** of anything contains 6.022×10^{23} objects. For instance, one mole of eggs contains 6.022×10^{23} eggs, one mole of carbon contains 6.022×10^{23} carbon atoms, and one mole of carbon dioxide contains 6.022×10^{23} carbon dioxide molecules.

The symbol of moles is “*mol*”¹.

1. The plural of mole is moles, but when using it as a unit, there is no need to use mols (just mol; eg. 3 moles = 3 mol).

- Thus, three moles (3 mol) of particles is three times 6.022×10^{23} particles ($3 \times 6.022 \times 10^{23}$), and ten moles (10 mol) of particles is ten times 6.022×10^{23} particles



Figure 5.1.1: Amedeo Avogadro – the unit is named after his contributions to chemistry. Image attribution: Amadeo_Avogadro is licensed under a Public Domain license.

$(10 \times 6.022 \times 10^{23})$.

- If we had 3.011×10^{24} atoms of gold (Au), the number of moles of gold would be:

$$n(\text{Au}) = \frac{3.011 \times 10^{24} \text{ atoms}}{6.022 \times 10^{23} \text{ atoms mole}^{-1}} = 5.000 \text{ mol}$$

This exact number is a fundamental SI unit within science. Avogadro's number N_A is the term used for 6.022×10^{23} objects.

Avogadro's number is the link between the number of atoms (or molecules or ions) of a material and the number of moles of material. Understanding it allows chemical equations to be represented through the ratios of compounds and particles – enabling us to upscale reactions outside of individual atoms to notable masses and quantities for measurement and experimentation.

Key Takeaways

- 1 mole is considered to be 6.022×10^{23} atoms or molecules (Avogadro's number).
- Utilising moles as a unit is important to allow us to talk about specific atomic quantities easily.

Exercises

Practice Questions



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5.2 Molar Mass and the Stoichiometry

Learning Objectives

- Apply the use of molar mass to allow for measurable quantities of atoms and molecules through weight.
- Understand how the molar mass of molecules is determined through the periodic table.
- Utilising stoichiometric coefficient, determine the masses of reactants needed to produce a desired amount of product, and vice versa.

In 2.5 Isotopes and Relative Abundance, we introduced the **Dalton** (Da) and the **atomic mass unit** (amu). These tell us the mass of individual atoms – which is roughly equal to the number of protons and neutrons.

Naturally, their small size makes it impractical for us to measure by the individual atom. As such, **Avogadro's number** was constructed with the mass of carbon-12 in mind (the same element used in the definition of Da and amu), allowing us to determine the amount of moles present in a material by weight. As such, practicing chemists hardly work with amu, but with **Molar mass**; the amount of grams required for 1 mole ($gmol^{-1}$) of a substance. It is defined in such a way that it is equivalent in number to Daltons.

In the case of carbon-12, with an atomic mass of 12.01 Da, we can

Periodic Table of the Elements

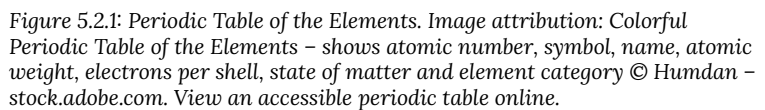


Table 5.2.1: Molar Masses of the First 20 Elements

Element	Molar Mass	Element	Molar Mass
H	1.01g mol^{-1}	Na	22.99g mol^{-1}
He	4.00g mol^{-1}	Mg	24.31g mol^{-1}
Li	6.94g mol^{-1}	Al	26.98g mol^{-1}
Be	9.01g mol^{-1}	Si	28.09g mol^{-1}
B	10.81g mol^{-1}	P	30.97g mol^{-1}
C	12.01g mol^{-1}	S	32.06g mol^{-1}
N	14.01g mol^{-1}	Cl	35.45g mol^{-1}
O	16.00g mol^{-1}	Ar	39.95g mol^{-1}
F	19.00g mol^{-1}	K	39.10g mol^{-1}
Ne	20.18g mol^{-1}	Ca	40.08g mol^{-1}

Molar Mass of Compounds

To calculate the molar mass of a compound or molecule, we must add the molar masses of each atom found within it.

Example: What is the molar mass of Sodium Carbonate?

The atomic mass of O, C and Na can be obtained from the periodic table in units of a gram: $O = 16.0\text{g/mol}$, $C = 12.0\text{g/mol}$, $Na = 23.0\text{g/mol}$. Sodium carbonate contains two Na atoms, one C atom and three O atoms:

$$\begin{aligned} \text{The molar mass} &= \text{Sum of masses of} \\ \text{atoms in a mole of the chemical species} \\ &= (2 \times 23.0 \text{ g/mol}) + 12.0 \text{ g/mol} \\ &+ (3 \times 16.0 \text{ g/mol}) \\ &= 106 \text{ g/mol} \end{aligned}$$

In sodium carbonate, 3 atoms of oxygen are needed to form only 1 molecule. We can directly relate this to the 3 moles of oxygen needed to form 1 mole of sodium carbonate.

The definition of molar mass creates an important relationship between the amount of substance (moles) and the mass of the substance (m):

$$n = \frac{m}{M}$$

Where, M is molar mass (g mol^{-1})

m is the mass of the substance (g)

n = is the number of moles (mol)

This is one of the most important formulas in chemistry! It allows us to determine the amount of moles of a substance present within a reaction through its mass.

*Example: calculate the amount of silver present in 50.00g of silver.
The molar mass of Ag is 107.87 grams per mole.*

The question states to calculate the amount of silver, which means the number of moles of silver. The molar mass of silver and the mass of silver are given in the question.

Therefore, you can use: $n = \frac{m}{M}$.

```
\[\begin{align*}n &= \frac{m}{M} \\ n &= \frac{50.00\text{g}}{107.87\text{g/mol}} \\ n &= 0.4635\text{mol} \\\end{align*}\]
```

The Mole and Chemical Formulas

As we have learnt, chemical formulas indicate the number of atoms of each element present in a compound. For instance, in Na_2CO_3 , one molecule of sodium carbonate contains two atoms of sodium, one atom of carbon and three atoms of oxygen. Similarly, chemical formulas can also indicate the number of moles of atoms of each type of element present in one mole of a compound. Let's take the same example Na_2CO_3 again; this time, the chemical formula of sodium carbonate indicates that 2 moles of sodium, one mole of carbon and three moles of oxygen are present in one mole of sodium carbonate.

Example: The formula for aspirin is $\text{C}_9\text{H}_8\text{O}_4$. Calculate the mass of aspirin in grams present in 0.65mol of aspirin. The molar masses of **C**, **H** and **O** are:
 $\text{C} = 12.01\text{g/mol}$, $\text{H} = 1.01\text{g/mol}$, $\text{O} = 16.00\text{g/mol}$

The first step is to calculate the molar mass of aspirin from the given atomic molar masses:

$$\left(9 \times 12.01 \frac{\text{g}}{\text{mol}}\right) + \left(8 \times 1.01 \frac{\text{g}}{\text{mol}}\right) + \left(4 \times 16.00 \frac{\text{g}}{\text{mol}}\right) = 180.1 \frac{\text{g}}{\text{mol}}$$

The next step is to rearrange the molar mass equation to make m the subject:

$$M = \frac{m}{n} \\ m = M \times n$$

The final step is to substitute the given values with units to the equation and solve the equation:

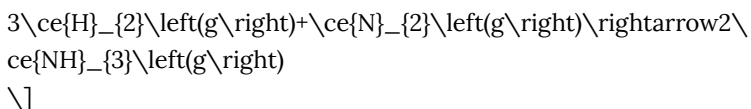
$$\begin{aligned} m &= M \times n \\ m &= 180.1 \frac{\text{g}}{\text{mol}} \times 0.65 \frac{\text{mol}}{\text{mol}} \\ m &= 117 \text{g} \end{aligned}$$

The relationship between mass, moles, and Avogadro's number is a concept known as **stoichiometry**. Stoichiometry is present throughout chemistry, allowing us to determine the values of molecules and compounds through mathematical equations and calculations.

Stoichiometric Coefficient

In chemical equations, the stoichiometric coefficient is the number written immediately before the atoms, ions and molecules in a balanced chemical equation. These numbers indicate the ratio of any reactant or product that reacts or forms in a reaction.

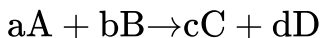
For example, let's consider the reaction between hydrogen gas and nitrogen gas to form ammonia. The balanced chemical equation can be written as follows:



Stoichiometric coefficients are the numbers found in front of the reactants and products of our chemical equation. 3 and 2 are written in front of H_2 , and NH_3 and are our coefficients. For components without a number, such as N_2 , we can treat this as a 1. Stoichiometric coefficients in this reaction show that three molecules of hydrogen gas react with one molecule of nitrogen gas to form two ammonia molecules.

The Molar Ratio and Chemical Equations

The stoichiometric coefficients in a balanced chemical equation can also be interpreted on the molar scale. For a general reaction such as:



the number of moles of A divided by its stoichiometric coefficient a is equal to the number of moles of B divided by its stoichiometric coefficient b , etc., as shown in below equation:

$$\frac{A_{\text{mol}}}{a} = \frac{B_{\text{mol}}}{b} = \frac{C_{\text{mol}}}{c} = \frac{D_{\text{mol}}}{d}$$

Let's consider the same example again:
$$3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$$

Based on the molar scale, we can say that three moles of hydrogen gas react with one mole of nitrogen gas to form two moles of ammonia gas. Molar scale interpretation relates to the quantities of substances involved in the reaction. Therefore, if we know the amount of one substance involved in the reaction, we can use the molar ratio to determine the amounts of others involved in the reaction.

*Example: How many moles of ammonia gas can be produced from **5mol** of hydrogen gas?*

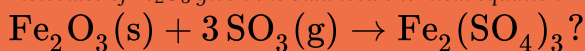
$$3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$$

We are asked to calculate the amount of ammonia and are given the moles of hydrogen and the balanced chemical equation for the reaction. We need to find out the molar ratio of hydrogen and ammonia to calculate the amount of ammonia. From the stoichiometric coefficients, we know that three moles of hydrogen produce two moles of ammonia. Therefore, the molar ratio of hydrogen gas to ammonia gas is **3 : 2**. Thus, if we use **5mol** of hydrogen gas, we can produce the following amount of ammonia gas:

$$\frac{2\text{mol NH}_3}{3\text{mol H}_2} \times 5$$

$$\text{mol H}_2 = 7.5 \text{ mol}$$

Example: How many molecules of SO_3 are needed to react with 144 molecules of Fe_2O_3 given the balanced chemical equation



Solution

We are asked to calculate the amount of sulphur trioxide and are given the moles of iron(III) oxide and the balanced chemical equation for the reaction. We need to find out the molar ratio of sulphur trioxide and iron(III) oxide. From the stoichiometric coefficients, we know that one mole of iron(III) oxide reacts with three moles of sulphur trioxide.

Therefore, the molar ratio is **1 : 3**. Thus, if we use

144 mol of iron(III) oxide, we need the following amount of sulphur trioxide:

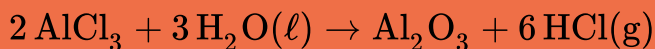
$$\frac{3 \text{ mol SO}_3}{1 \text{ mol Fe}_2(\text{SO}_4)_3} \times 144 \text{ mol Fe}_2(\text{SO}_4)_3 = 432 \text{ mol}$$

We need 432 moles of SO_3 to react with 144 moles of Fe_2O_3 .

Mass Calculations

Now knowing how many moles are necessary of certain compounds, we can determine the weight of reactants needed to perform a reaction. Using the periodic table, we can calculate the molar masses of the reactants by adding individual atomic weights.

Example: How many moles of HCl will be produced when 249 g of AlCl_3 are reacted according to this chemical equation?



We will do this in two steps: convert the mass of AlCl_3 to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of AlCl_3 is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$n = \frac{m}{M}$$
$$\frac{249\text{gAlCl}_3}{133.33\text{gmol}^{-1}\text{AlCl}_3} = 1.87\text{mol}$$

$$1.87 \cancel{\text{ mol AlCl}_3} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{ mol AlCl}_3}} = 5.61 \text{ mol HCl}$$

Now, we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of AlCl₃ and the number of moles of HCl:

$$\frac{6 \text{ mol HCl}}{2 \text{ mol AlCl}_3}$$

Applying this conversion factor to the quantity of AlCl₃, we get:

$$1.87 \cancel{\text{ mol AlCl}_3} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{ mol AlCl}_3}} = 5.61 \text{ mol HCl}$$

Alternatively, we could have done this in one line:

$$249 \cancel{\text{ g AlCl}_3} \times \frac{1 \cancel{\text{ mol AlCl}_3}}{133.33 \cancel{\text{ g AlCl}_3}} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{ mol AlCl}_3}} = 5.60 \text{ mol HCl}$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

Stoichiometry is one of the fundamental concepts of chemical sciences. It is highly recommended that students get familiar and acquainted with the mathematical approaches employed in this and the next chapter, as future study will assume students' knowledge in this area.

Key Takeaways

- Molar mass allows us to convert the mass of a substance present to the number of moles, using the atomic weight of elements from the periodic table.
- Molar mass of compounds is determined through adding the molar masses of individual elements.
- The moles and therefore mass of any products and reactants within a reaction can be determined through the stoichiometric coefficient.

Exercises

Practice Questions



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Transcript

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- Colorful Periodic Table of the Elements – shows atomic number, symbol, name, atomic weight, electrons per shell, state of matter and element category © Humdan - stock.adobe.com is licensed under a All Rights Reserved license

5.3 Concentration and Solutions

Learning Objectives

- Define minor and major components within solutions: solutes and solvents.
- Understand how to represent concentrations through a variety of units: percentage concentration, parts per million, parts per billion, and molarity.
- Use molarity calculations to determine the mass of solute needed to create solutions of specific concentrations.

Chemistry often involves the use of solutions to facilitate reactions. In a liquid solution, a compound (of any **state of matter**) is dissolved into a liquid medium. The major component of a solution is called the **solvent**. The minor component of a solution is called the **solute**. By major and minor we mean whichever component has the greater presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, here we will confine the discussion to solutions for which the major component and the minor component are obvious. As a general rule of thumb for liquid solutions, water, alcohol and other liquids in large quantities are the **solvents** – while dissolved salts, acids, or anything added to a liquid is the **solute**.

Solutions exist for every possible state of matter combination of

the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water; soda water is a solution of gaseous CO₂ in liquid water, while air is a solution of a gaseous solute (O₂) in a gaseous solvent (N₂). In all cases, however, the overall **phase** of the solution is the same **phase** as the solvent.

Example: A solution is made by dissolving 1.00 g of sucrose (C₁₂H₂₂O₁₁) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Either by mass or by moles, the obvious minor component is sucrose, so it is the solute. Water — the majority component — is the solvent. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Water is a polar solvent that is commonly used throughout chemistry thanks to its easy-to-reach boiling and melting points, alongside availability. It is so common that for chemical equations, a special subscript is used to quickly denote that a solvent is dissolved in a water solute: (aq). This is referred to as ‘aqueous’. Commonly, acids (such as hydrochloric or nitric) exist as solutes in water solvents. As such, they are commonly shown as $HCl_{(aq)}$ or $HNO_{3(aq)}$ to showcase this

fact, removing the need to mention water unless it is directly involved in the reaction.

On a technical note, H_2O can never be considered aqueous – by definition, water cannot be dissolved by water. It is always referred to as a liquid (denoted by (l)). We will however let you decide whether water is wet!

Concentration and Dilution of Solutions

One important concept of solutions is in defining how much solute is dissolved in a given amount of solvent. This concept is called **concentration**. Various words are used to describe the relative amounts of solute. **Dilute** describes a solution that has very little solute, while **concentrated** describes a solution that has a lot of solute. One problem is that these terms are qualitative; they describe more or less but not an exact amount. There are a variety of ways to show a quantitative description of the concentration of a solution. Different methods of expressing concentration include:

1. Percent concentration (%)
2. Parts per million (ppm)

3. Parts per billion (ppb)
4. Molarity (M)

Percent concentration

Percentage concentrations are useful to show the direct comparison between a solvent and a solute. There are three ways of expressing concentration as a percentage.

1. Mass/mass percent concentration, (m/m)% $\left[\frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100 \right]$

*Example: What is the percent by mass concentration of sodium chloride solution prepared by dissolving **3.25g** of sodium chloride in **42.50g** of water?*

Masses of the solute and the solvent are given, and we are asked to calculate the percent concentration of sodium chloride by mass:

$$= \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100$$

$$= \frac{3.25\text{g}}{42.50\text{g} + 3.25\text{g}} \times 100$$

$$= 7.1\%$$

2. Volume/volume percent concentration, (v/v)% $\left[\frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100 \right]$

$$\text{right})\}\frac{\text{volume of solute}}{\text{volume of solution}}\times 100\%$$

Example: If 8.00mL of ethanol is dissolved in enough water to give 80.00mL of the solution, what is the percent by volume concentration of ethanol in the resulting solution?

We are asked to calculate ethanol concentration in percent by volume. The volumes of the solute and the solution are given:

$$\begin{aligned} &= \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \\ &= \frac{8.00\text{mL}}{80.00\text{mL}} \times 100\% \\ &= 10.0\% \end{aligned}$$

3. Mass/volume percent concentration, (m/v)%
- $$\frac{\text{mass of solute}}{\text{volume of solution}} \times 100\%$$

Example: If 5.0g of MgCl_2 is dissolved in enough water to give 200mL of solution, calculate the concentration of MgCl_2 in (m/v)%.

We are asked to calculate the concentration of MgCl_2 in (m/v)%. The mass of the MgCl_2 and the volume of the resulting solution are given:

$$\begin{aligned} &= \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100 \\ &= \frac{5.0 \text{ g}}{200 \text{ mL}} \times 100 \\ &= 2.5\% \end{aligned}$$

Parts Per Million and Parts Per Billion

Parts per million (ppm) and parts per billion (ppb) are typically used to describe trace quantities of solutes found within a solution. They represent the parts of an analyte found within one million or one billion parts of a medium respectively.

Parts can be in either mass or volume units. Both solute and solvent must have the same units. For example, one gram of solute is present in a million grams of solution, and one millilitre of benzene is present in a million millilitres of air. To illustrate:

$$1 \text{ ppm} = \frac{1 \text{ mL}}{1000000 \text{ mL}} \times 1000000$$

Commonly, however, analytes are found dissolved in water. As the density of water is extremely close to 1 g mL^{-1} , it is often appropriate to report 1 ppm as equal to 1 mg L^{-1} for aqueous solutions near room temperature.

Example: Drinking water contains 0.002g of Mg^{+} ions in 100g of water. What is this concentration in parts per million?

We are asked to calculate the given concentration in ppm:

$$= \frac{0.002\text{g}}{100\text{g}} \times 1000000$$

$$= 20\text{ppm}$$

Molarity (M)

Molarity is the number of solute moles present in one litre of a solution. This is the most common way to show the concentration of a solution within chemical laboratories thanks to its direct relationship to moles (which, as discussed in 5.2 'Molar Mass', can be easily determined by mass). We can determine molarity by rearranging the following formula:

$$n = cv$$

Where, n is the number of moles (mol)
 c is the molarity of the solution (mol L^{-1})
 v is the volume of the solution (L)

This is one of the most important formulas in chemistry! It allows us to determine the amount of moles of a substance present within a given concentration and the volume of a substance. However, rearranging allows us to determine the moles required

to create a desired concentration for a given volume. As such, a common version of the above equation utilised by chemists is as follows:

$$c = \frac{n}{v}$$

Example: If 2.50mol of CuSO_4 are dissolved in enough water to give a 500.0mL solution, what is the molarity of the resulting solution?

We are given the number of moles present in 500mL of a solution and asked to calculate the molarity. Molarity means the number of moles present in 1000mL or 1L of solution. Recall that $c = \frac{n}{v}$. Therefore:

$$\begin{aligned} &= \frac{\text{moles of solute}}{\text{volume of solution}} \\ &= \frac{2.50\text{mol}}{0.5\text{L}} \\ &= 5.0\text{M} \end{aligned}$$

Have a look at the interactive simulation (Figure 5.3.1) to observe how moles and volume interact with each other to determine molarity:





One or more interactive elements has been excluded from this version of the text. You can view them online here: <https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=916>

Figure 5.3.1: Simulation by PhET Interactive Simulations, Univer

In combination with $n = \frac{m}{M}$, these equations allow chemists to determine the mass of a solute needed to achieve a specific concentration within a solution. These two equations represent the majority of undergraduate chemistry calculations – try to remember both off by heart! Let's observe how we can find the mass of a compound needed to achieve a desired concentration.

Example: How many grams of glucose are needed to make a **250mL** of a **1.5M** glucose solution? The molar mass of glucose is **180.15g/mol**.

Here, we are given the molarity of the solution and asked to calculate the mass of the solute present in a given volume. Let's utilise $n = CV$ to determine the moles needed for the solution. Then, we will use the $n = \frac{m}{M}$ equation to convert moles to grams.

Number of moles of glucose present in **250mL** solution:
$$\begin{aligned} & \end{aligned}$$

$$= 1.50 \text{ mol/L} \times 0.250 \text{ L}$$

$$= 0.375 \text{ mol}$$

Grams of glucose present in **250ml** solution:

$$n = \frac{m}{M}$$

$$n \times M = m$$

$$0.375 \text{ mol} \times 180.15 \text{ g/mol} = m$$

$$m = 67.6 \text{ g}$$

*Example: How many millilitres of a **1.00M HCl** solution would be required to obtain **7.3g** of **HCl**? The molar mass of **HCl** is **36.458g/mol**.*

We are given the molarity of the solution and asked to calculate the volume that the given grams of the solute would present. First, convert the grams of the solute into moles using $n = \frac{m}{M}$. Then, molarity is used as the conversion factor to find the volume that the calculated moles would present.

required to produce 1.0 mol/L using 0.20 mol of HCl.”

Number of moles of **HCl** corresponds to $n = \frac{m}{M}$,

where, n = number of moles m = mass of the substance M = Molar mass of the substance:

$$\begin{aligned} n &= \frac{m}{M} \\ n &= \frac{7.3 \text{ g}}{36.458 \text{ g/mol}} \\ n &= 0.20 \text{ mol} \end{aligned}$$

The volume that required to produce 1.0 mol/L using 0.20 mol of HCl :

$$\frac{1000 \text{ mL}}{1.00 \text{ mol}} \times 0.20 \text{ mol} = 200 \text{ mL}$$

Dilution of Solutions

Dilution is a process used to lower the concentration of the original solution by adding solvents. The concentrated solution is known as the stock solution. However, the amount of solute present in the solution remains constant before and after the dilution; only the volume changes. The following equation shows the relationship between concentration and volume before and after dilution. This equation is built upon the fact that the same amount of solute is present in the solution before (stock solution) and after (diluted solution) the dilution:

$$C_1 V_1 = C_2 V_2$$

Where,

C_1 = Concentration of the stock solution

V_1 = Volume of the stock solution

C_2 = Concentration of the diluted solution

V_2 = Volume of the diluted solution

*Example: A chemist prepared a **10mL** of **1.00M HCl** solution by adding water to the **3.0M** stock solution of **HCl**. How much stock solution should be taken for the dilution?*

We are given the concentration of the diluted solution, volume of the diluted solution, and concentration of the stock solution and asked to calculate the volume of the stock solution used for the dilution. As the number of solute moles remains constant before and after the dilution, we can use the $C_1 V_1 = C_2 V_2$ equation to calculate the volume of the stock solution:

$$\begin{aligned} C_1 V_1 &= C_2 V_2 \\ 3.0 \text{ M} \times V_1 &= 1.0 \text{ M} \times 10 \text{ mL} \\ V_1 &= \frac{1.0 \text{ M} \times 10 \text{ mL}}{3.0 \text{ M}} \\ V_1 &= 3.3 \text{ mL} \end{aligned}$$

Chemistry Is Everywhere: Preparing IV Solutions

In a hospital emergency room, a physician orders an intravenous (IV) delivery of 100 mL of 0.5% KCl for a patient suffering from hypokalemia (low potassium levels). Does an aide run to a supply cabinet and take out an IV bag containing this concentration of KCl (Figure 5.3.2)?

Not likely. It is more probable that the aide must make the proper solution from an IV bag of sterile solution and a more concentrated, sterile solution, called a *stock solution*, of KCl. The aide is expected to use a syringe to draw up some stock solution and inject it into the waiting IV bag and dilute it to the proper concentration. Thus the aide must perform a dilution calculation.



Figure 5.3.2: Medical personnel commonly must perform dilutions for IV solutions. Image attribution: “Infuuszakjes” © 2008 by Harmid is licensed under a Public Domain license.

If the stock solution is 10.0% KCl and the final volume and concentration need to be 100 mL and 0.50%, respectively, then to determine how much stock solution to use:

$$(10\%)V_1 = (0.50\%)(100 \text{ mL})$$

$$V_1 = 5 \text{ mL}$$

Of course, the addition of the stock solution affects the total volume of the diluted solution, but the final concentration is likely close enough, even for medical purposes.

Medical and pharmaceutical personnel are constantly dealing with dosages that require concentration measurements and dilutions. It is an important responsibility: calculating the *wrong* dose can be useless, harmful, or even fatal!

In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table 5.3.1 “Solubilities of Some Ionic Compounds” lists the solubilities of some simple ionic compounds. These solubilities vary widely: NaCl can dissolve up to 31.6 g per 100 g of H₂O, while AgCl can dissolve only 0.00019 g per 100 g of H₂O.

Table 5.3.1 Solubilities of Some Ionic Compounds.

Solute	Solubility (g per 100 g of H ₂ O at 25°C)
AgCl	0.00019
CaCO ₃	0.0006
KBr	70.7
NaCl	36.1
NaNO ₃	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solvent, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H₂O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H₂O is also saturated but rather concentrated. Ideally, we need more precise ways of specifying the amount of solute in a solution.

In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.

It should be obvious that some solutes dissolve in certain solvents but not others. NaCl, for example, dissolves in water but not in vegetable oil. Beeswax dissolves in liquid hexane but not water. What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. Intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding. From experimental studies, it has been determined that if the molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl — a very polar substance because it is composed of ions — dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane but not in polar water. This concept leads to the general rule that “**like dissolves like**” for predicting whether a solute is soluble in a given solvent (see Figure 5.3.3).

However, this is a general rule, not an absolute statement, so it must be applied with care.

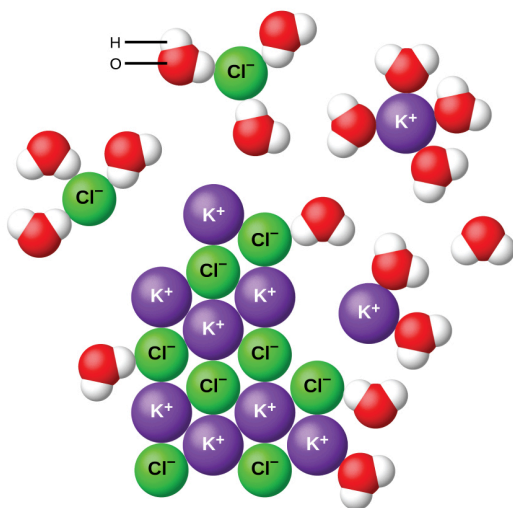


Figure 5.3.3: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown. Transcript. Image attribution: OpenStax: licensed under a Creative Commons Attribution 4.0 International License.

Example: Would I_2 be more soluble in CCl_4 or H_2O ? Explain your answer.

Solution

I_2 is a nonpolar molecule due to the similar electronegativities of the iodine atoms.

CCl_4 is a nonpolar molecule: In CCl_4 , electrons are more attracted towards the chlorine atoms due to the high electronegativity of chlorine compared to the carbon. However, as four bonds are arranged symmetrically, the dipole moment of one bond is cancelled out by the opposite bond, creating a zero net dipolar moment. Therefore, CCl_4 , is nonpolar. View this video to learn why CCl_4 , is a nonpolar molecule.

H_2O is a polar molecule: In the water molecule, due to the high electronegativity of the oxygen atom, electrons are strongly attracted towards the oxygen atom. Additionally, the water molecule has a bent, or angular, structure due to the two lone pairs of electrons on the oxygen atom. This arrangement results in an asymmetrical charge distribution within the molecule, with the oxygen end being more negative and the hydrogen end being more positive. Therefore, the combination of the polar covalent bonds and the molecular geometry leads to the water molecule having a net dipole moment, making it a polar molecule. View this video to learn why H_2O , is a polar molecule.

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar, and H_2O is polar. Since “like dissolves like,” nonpolar solutes tend to dissolve better in nonpolar solvents. So I_2 would be expected to be more soluble in CCl_4 .

Key Takeaways


- Solutions contain at least two components: a large amount of solvent and a smaller amount of solute(s)
- Species dissolved in water are called “aqueous solutions”.
- There are four common measures of concentration: percentage concentration, parts per million, parts per billion, and molarity.
- In chemistry, “like-dissolves-like”. Polar solvents will generally dissolve polar solutes, while non-polar solvents dissolve non-polar solutes.

Exercises

Practice Questions



An interactive H5P element has been excluded

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Transcript

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5.4 Limiting Reagents and Yield

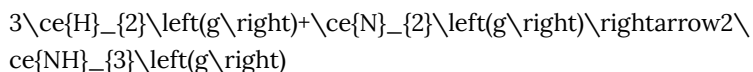
Learning Objectives

- Utilising stoichiometry coefficients and molar mass calculations, determining limiting reagents.
- Applying limiting reagents to calculate theoretical yields – allowing for determination of overall yield.

So far, we are yet to consider some of the practical limitations encountered when considering real chemical reactions in the laboratory. We have assumed that reactions proceed all the way to completion and that all reactants are **in excess**. In the real world, this is not always the case. In this chapter, let's address those assumptions.

Limiting and Excess Reagents

Let's consider the reaction between hydrogen gas and nitrogen gas to form ammonia gas: \[



\]

Based on stoichiometry, we know that if we mix three moles of

H_2 with one mole of N_2 we can obtain two moles of NH_3 . In other words, H_2 reacts with N_2 in a molar ratio of **3 : 1**. If we mix three moles of hydrogen gas and two moles of nitrogen gas, what will happen? Similarly, three moles of H_2 will react with one mole of N_2 and produce two moles of NH_3 . However, the entire quantity of H_2 will be used before the N_2 as we start with the exact amount of H_2 required for the reaction. The product mixture will contain one unreacted mole of N_2 as we only required one mole of N_2 for the reaction. If we had more H_2 we could have produced more NH_3 as we already have excess N_2 . Therefore, in this reaction, we call H_2 the limiting reagent since it limits the production of NH_3 .

The limiting reagent is the reactant that runs out first, limiting the formation of products. In the above example, nitrogen gas is called the excess reagent because we have a greater amount than required for the reaction.

Example: phosphoric acid is produced by treating phosphate rock, which contains calcium phosphate, with sulphuric acid as follows:

$$3\text{Ca}_3(\text{PO}_4)_2 + 3\text{H}_2\text{SO}_4 \rightarrow 3\text{CaSO}_4 + 2\text{H}_3\text{PO}_4$$

A chemist in a fertiliser production plant mixed **144.5g** of calcium phosphate with **160.5g** of sulphuric acid. What is the maximum mass of phosphoric acid (H_3PO_4) that can be formed? Molar masses are given in table 5.4.1.

Table 5.4.1: Molar masses for production of phosphoric acid

Molecule	Molar mass
$\text{Ca}_3(\text{PO}_4)_2$	310.2
H_2SO_4	98.1
H_3PO_4	97.9

We are given the mass of reactants and are asked to determine the mass of one of the products.

Step 1 – Identify the limiting reagent in this reaction. For that, find the number of moles of each reactant using

$$n = \frac{m}{M}.$$

Number of moles of $\text{Ca}_3(\text{PO}_4)_2$:

$$n = \frac{m}{M}$$

$$n = \frac{144.5 \text{ g}}{310.2 \text{ g/mol}}$$

$$n = 0.4658 \text{ mol}$$

Number of moles of H_2SO_4 :

$$n = \frac{m}{M}$$

$$n = \frac{160.5 \text{ g}}{98.1 \text{ g/mol}}$$

$$n = 1.64 \text{ mol}$$

Then, compare the calculated moles with the molar ratio of the two reactants. According to the balanced chemical equation, calcium phosphate reacts with sulphuric acid in a **1 : 3** molar ratio. Therefore, **0.4658mol** of

$\text{Ca}_3(\text{PO}_4)_2$ reacts with the following number of moles of H_2SO_4 :

$$\frac{3 \text{ mol H}_2\text{SO}_4}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \times 0.4658 \text{ mol Ca}_3(\text{PO}_4)_2 = 1.397 \text{ mol}$$

We have 1.64 mol of H_2SO_4 , but we only need 1.397 mol for the reaction. This means we have excess H_2SO_4 acid. Therefore, the limiting reagent is $\text{Ca}_3(\text{PO}_4)_2$. The quantity of limiting reagent can determine the mass of H_3PO_4 acid formed.

Step 2 – Use the molar ratio between $\text{Ca}_3(\text{PO}_4)_2$ and H_3PO_4 acid to calculate the mass of H_3PO_4 acid formed.

We know one mole of $\text{Ca}_3(\text{PO}_4)_2$ forms two moles of H_3PO_4 acid from the balanced chemical equation. Therefore, the number of moles of H_3PO_4 acid forms from 0.4658 mol of $\text{Ca}_3(\text{PO}_4)_2$ is:

$$\frac{2 \text{ mol H}_3\text{PO}_4}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \times 0.4658 \text{ mol Ca}_3(\text{PO}_4)_2 = 0.9316 \text{ mol}$$

Now, find the mass of H_3PO_4 using $M = \frac{m}{n}$


```

n & =\frac{m}{M}\\
m & =M\times n\\
m & =97.99\text{g}/\text{mol}\times0.9316\\
& \text{mol}\\
m & =91.29\text{g}\\
\end{align*}

```

Based on the given reaction conditions, the maximum mass we can obtain for H_3PO_4 acid is **91.29g**.

One of the fundamental principles in chemistry and physics is the **law of conservation of energy and mass**. This famous rule explains that we cannot create nor destroy matter, only transform it. Just like baking a cake, you can only get as much as you put it! The amount of reactants present is our core limitation of how much a reaction can produce – which will form a foundational concept within the topic of yields.

Yields

As we have just seen, stoichiometry allows us to determine the maximum mass or amount of a product a reaction can generate for a given quantity of reactants. In reality, chemists never extract the full 100% of a generated product in a reaction.

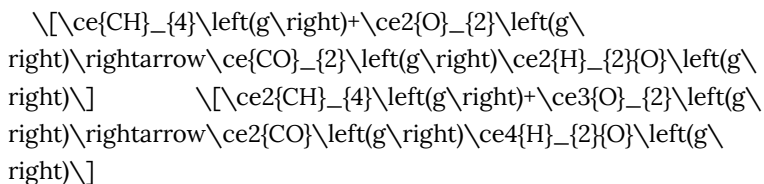


Figure 5.4.1: Notice how the brown staining on the lip of the mug presents. Within it contains caffeine yet to be recovered! Image attribution: A cup of coffee with heart shaped latte design presented on a wooden surface in the morning. © asiraj – stock.adobe.com.

Consider a mug of coffee. Once you have ingested the liquid, there is often a visible brown stain in the mug due to the adhesive forces of the liquid (see Figure 5.4.1). There is likely to be trace amounts of caffeine that you have yet to digest. While you could fill the cup with water and attempt to recover the caffeine, eliminating the visible colour, on the molecular scale with

sextillions of molecules, it is impossible to capture every compound. Even in perfect reactions that go to completion with many repetitive washes to gather all of a produced product, we are unable to confidently have recovered everything.

From a more chemical perspective, however, side reactions are a common occurrence in many processes. For example, in the combustion of some carbon-containing compounds, such as methane, there are 2 potential reactions that can occur depending on the available levels of oxygen. They are:



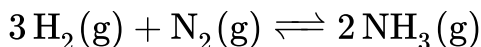
For now, don't worry about side reactions

Complete combustion (with oxygen excess) generates CO_2 (carbon dioxide), whereas incomplete combustion produces CO (carbon monoxide) – a more toxic greenhouse gas responsible for smog! Car manufacturers attempt to design engines with sufficient airflow to

allow for excess oxygen in the environment. However, as they might, some CO is always produced. This, alongside every other chemical reaction, is inefficient to some degree. It is impossible to recover 100% of a reaction.

and
equilibriums. We
will bring them
up if important.

Finally, we can address our first assumption, that all reactions go to completion. This is not true. While most reactions will undergo enough conversion that we can realistically say they have completed fully, some reactions instead exist in equilibrium, where only a certain amount of potential product is produced. The equation we looked at the beginning of this chapter, the production of ammonia, is one of the most famous equilibrium reactions, known as the Haber process:



In the Haber process, reactants and products exist in equilibrium with each other (denoted by the two-way arrow in the equation). This equilibrium can be influenced by air pressure and temperature – which, if carefully managed, allows for the extraction of ammonia. However, this system will never allow for the full 100% extraction of the gas.

This is all to say that in chemical processes, we must consider our **percentage yield**. Percentage yield indicates the success of the reaction. If the value is closer to 100%, that means losses are minor, and most reactants have converted to products. This is calculated by considering the theoretical yield. as previously demonstrated, and the actual yield, measured in a laboratory, in the following formula:

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

The theoretical yield is the maximum amount of the product that

can form under the given reaction conditions. The available amount of limiting reagent determines the maximum amount of the product that can be formed. As demonstrated previously, theoretical yield can be calculated using the stoichiometric coefficients (or molar ratios) displayed in the balanced chemical equation. We calculate theoretical yield assuming that no losses occur during the reaction and all the reactants convert to products. However, in reality, losses and inefficiencies often occur during chemical reactions. Therefore, the actual amount that you isolate and measure from a chemical reaction is less than the theoretically expected amount.

The actual yield of a product is the amount of the product you isolate from the chemical reaction. You cannot calculate the actual yield. It is required to be determined experimentally, generally through weighing the product produced.

From these two values, the percentage yield can be calculated. As a general rule of thumb, yield decreases with the number of steps needed to complete a chemical reaction but can be increased through repeat filtrations and extractions within an experiment. In chemistry and especially industrial applications, bolstering percentage yield is essential to cost-effective operations and the adoption of processes.

Example: When 168.3g of Fe react with O₂ according to the following chemical equation, 172.6g of Fe₂O₃ is obtained.

What is the percentage yield of Fe₂O₃ for this reaction? The molar mass of Fe₂O₃ is 159.69g/mol and Fe is

55.85g/mol. \[



We are given the mass of the reactant **Fe**, the actual yield of the product **Fe₂O₃**, the balanced chemical equation and are asked to calculate the percentage yield of **Fe₂O₃** for this reaction. We need to find out the theoretical yield of **Fe₂O₃** for this reaction to calculate the percentage yield.

Step 1 – Use the molar ratios given in the balanced chemical equation and the given mass of **Fe** to calculate the theoretical yield of **Fe₂O₃**:

$$\begin{aligned} \text{Number of moles of Fe: } n &= \frac{m}{M} \\ n &= \frac{168.3 \text{ g}}{55.85 \text{ g/mol}} \\ n &= 3.013 \text{ mol} \end{aligned}$$

From the balanced chemical equation, we know four moles of **Fe** form two moles of **Fe₂O₃**. Therefore, **3.013 mol** of **Fe** forms the following amount of **Fe₂O₃**:

$$\frac{2 \text{ mol}}{4 \text{ mol}} \times 3.013 \text{ mol} = 1.507 \text{ mol}$$

Convert the number of moles to **Fe₂O₃**

$$\begin{aligned} \text{mass: } m &= M \times n \\ m &= 159.69 \text{ g/mol} \times 1.507 \text{ mol} \\ m &= 240.7 \text{ g} \end{aligned}$$

Step 2 – Calculate the percentage yield of Fe_2O_3 using the calculated theoretical yield and the given actual yield:

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{172.6 \text{ g}}{240.7 \text{ g}} \times 100 = 71.71\%$$

Key Takeaways

- Reactions are limited by the reactant that runs out first – known as the limiting reagent. This is calculated by comparing the molar ratios against the moles of reactant actually used in the experiment for each substance in a reaction.
- Theoretical yield is the maximum amount of product that can be generated in a chemical reaction according to the limiting reagent.
- Experimental yield is the amount of product actually recovered.
- Percentage yield is the fraction of the experimental yield compared to the theoretical yield.

Practice Questions



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<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=1447#h5p-70>

Transcript

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- A cup of coffee with heart shaped latte design presented on a wooden surface in the morning. © asiraj - stock.adobe.com is licensed under a All Rights Reserved license

5.5 Standard Lab Conditions and the Ideal Gas Law

Learning Objectives

- Gain an appreciation of the need for replicability within scientific investigations.
- Define standard laboratory conditions (SLC) and the expected variance between experimental locations.

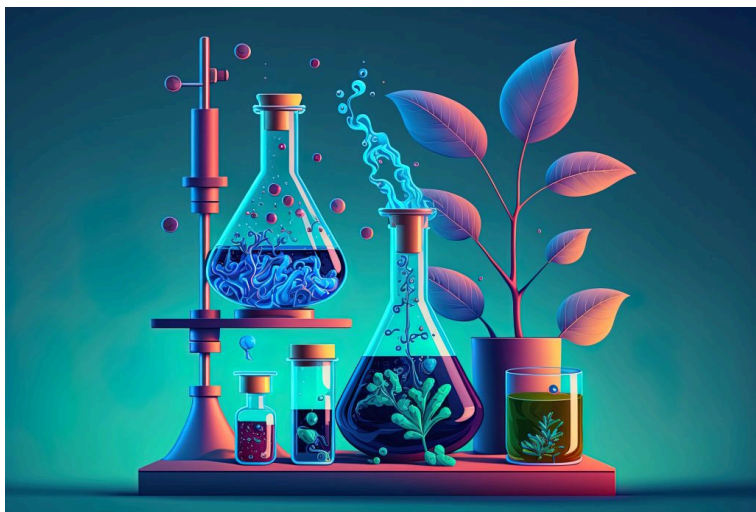


Image attribution: Blue backdrop with a chemistry lab set on a table with colourful substances within. Glassware and biological apparatus. Generative AI © 2ragon – stock.adobe.com.

The exact way that reactions process is highly dependent on the environmental conditions in which they occur. While the last chapter, 5.4 'Limiting Reagents and Yield', concerned the amount of reactants present, this understanding can also be applied to the forces and energy experienced by molecules. Temperature and pressure are two common factors that must be considered in chemical interactions. Let's consider an example of how they interact.

The Ideal Gas Law

The ideal gas law utilises a variety of observable gas properties to create one universal equation to model ideal gas behaviour. It is derived from the following laws:

- Boyle's Law: pressure and volume are inversely related at constant temperature and moles of gas.
- Charles's Law: volume and temperature are directly related at constant pressure and moles of gas.
- Guy-Lussac's Law: pressure and temperature are directly related at constant volume and moles of gas.
- Avogadro's Law: volume and moles are directly related at constant pressure and temperature.

As a result, a 'gas constant' can be produced by combining these laws together. Therefore:

$$\frac{PV}{nT} = \text{constant} = R$$

which is usually rearranged as follows:

$$PV = nRT$$

This equation is called the ideal gas law. It relates the four independent properties of a gas at any time. The constant R is called the ideal gas law constant. Its value depends on the units used to express pressure and volume (see table 5.5.1).

Table 5.5.1: Values of the Ideal Gas Law Constant R

Numerical Value	Units
0.08205	$\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$
62.36	$\frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}} = \frac{\text{L} \cdot \text{mmHg}}{\text{mol} \cdot \text{K}}$
8.314	$\frac{\text{J}}{\text{mol} \cdot \text{K}}$

This idealised law is not perfectly representative of real gases, as it does not take into account the intermolecular forces experienced by gas molecules. Instead, an alternative formula known as the van

der Waals equation is utilised to predict real behaviours. However, the idealised equation allows us to observe general trends.

Let us imagine an experiment conducted measuring the ambient pressure of an ideal gas in a sealed container at room temperature. On a cold night, a faulty laboratory in Antarctica that is still waiting for the heating to be repaired reports minimal amounts of pressure experienced by this gas at room temperature. They note that their measurements fluctuated greatly over the night and publish their findings in a well-respected journal. These findings did seem to contradict established gas laws; it would be a shock for the entire science world! With controversial results, a laboratory in Australia, more well equipped to maintain a constant room temperature of 25°C, performs the same experiment and invalidates their results, noting that Guy-Lussac's findings and the ideal gas law still stand. They recommend an available engineer to assist the Antarctic researchers in repairing the air-conditioning! It would seem that the environmental conditions in which a reaction occurs are of utmost importance.

Standard Laboratory Conditions

It is not just gas molecules that are influenced by laboratory conditions. Within chemistry, heat and pressure influence all reactions to some degree. As the reproducibility of experiments is needed to validate any findings, scientists must take note in their

writings and methodologies of factors such as temperature, pressure, humidity, or any other external factor which could alter the outcome of an experiment. Often, experiments are conducted without any change to these factors.

As such, the term **standard laboratory conditions** describes a typical, room-temperature laboratory. Standard laboratory conditions (SLC) are centred around 25°C (298.15 K) at 1 atm (101.325 KPa), experiencing 1g (9.81ms^{-2}) of gravity. The reason for this vagueness in the definition is due to geographical differences between laboratories. Aspects such as height from sea level (influencing pressure and temperature) and distance from the poles (due to Earth's imperfect shape presenting differences in gravity) will cause laboratories to experience differences in conditions. Generally, the differences between laboratories (that are well maintained) are so minor that they are negligible in standard experimentation. However, for work and research desiring extreme accuracy, these conditions can have notable impacts on outcomes.

Key Takeaways

- The ideal gas law describes relationships between P, V, T and n of a gas whose component molecules are assumed to have NO intermolecular interactions with each other (and not directly to do with 'ideal conditions').
- Most experiments are completed under standard laboratory conditions, with no extreme deviation from average conditions.

Practice Questions



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PART VI

CHAPTER 6: ACIDS, BASES AND SALTS

6.1 Definitions and Theories of Acids and Bases

Learning Objectives

- Identify an Arrhenius acid and an Arrhenius base.
- Write the chemical reaction between an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

Arrhenius Acids and Bases

Historically, the first chemical definition of an acid and a base was put forward by Svante Arrhenius, a Swedish chemist, in 1884. An **Arrhenius acid** is a compound that increases the H^+ ion concentration in aqueous solution. The H^+ ion is just a bare proton, and it is rather clear that bare protons are not floating around in an aqueous solution. Instead, chemistry has defined the **hydronium ion** H_3O^+ as the actual chemical species that represents an H^+ ion. H^+ ions and H_3O^+ ions are often considered interchangeable when writing chemical equations

(although a properly balanced chemical equation should also include the additional H_2O). Classic Arrhenius acids can be considered ionic compounds in which H^+ is the cation. Table 6.1.1 lists examples of Arrhenius acids and their names.

Table 6.1.1 Arrhenius Acids

Formula	Name
$\text{HC}_2\text{H}_3\text{O}_2$ (also written CH_3COOH)	acetic acid
HClO_3	chloric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid
HF	hydrofluoric acid
HNO_3	nitric acid
$\text{H}_2\text{C}_2\text{O}_4$	oxalic acid
HClO_4	perchloric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid
H_2SO_3	sulfurous acid

An **Arrhenius base** is a compound that increases the OH^- ion concentration in aqueous solution. Ionic compounds of the OH^- ion are classic Arrhenius bases.

Example 6.1.1

Problem

Identify each compound as an Arrhenius acid, an Arrhenius base, or neither.

1. HNO_3
2. CH_3OH
3. $\text{Mg}(\text{OH})_2$

Solution

1. This compound is an ionic compound between H^+ ions and NO_3^- ions, so it is an Arrhenius acid.
2. Although this formula has an OH in it, we do not recognize the remaining part of the molecule as a cation. It is neither an acid nor a base. (In fact, it is the formula for methanol, an organic compound.)
3. This formula also has an OH in it, but this time, we recognize that the magnesium is present as Mg^{2+} cations. As such, this is an ionic compound of the OH^- ion and is an Arrhenius base.

Test Yourself

Identify each compound as an Arrhenius acid, an Arrhenius base, or neither.

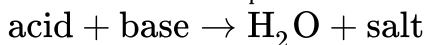
1. KOH
2. H_2SO_4
3. C_2H_6

Answer

1. Arrhenius base
2. Arrhenius acid
3. neither

Acids have some properties in common. They turn litmus, a plant extract, red. They react with some metals to give off H_2 gas. They react with carbonate and hydrogen carbonate salts to give off CO_2 gas. Acids that are ingested typically have a sour, sharp taste. (The name *acid* comes from the Latin word *acidus*, meaning “sour.”) Bases also have some properties in common. They are slippery to the touch, turn litmus blue, and have a bitter flavour if ingested.

Acids and bases have another property: they react with each other to make water and an ionic compound called a salt. A **salt**, in chemistry, is any ionic compound made by combining an acid with a base. A reaction between an acid and a base is called a **neutralisation reaction** and can be represented as follows:



The stoichiometry of the balanced chemical equation depends on

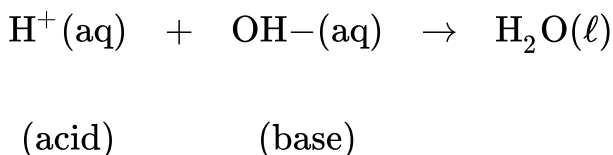
the number of H^+ ions in the acid and the number of OH^- ions in the base.

Brønsted-Lowry Acid-base Theory

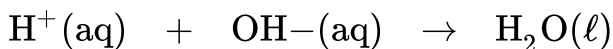
The Arrhenius definition of acid and base is limited to aqueous (that is, water) solutions. Although this is useful because water is a common solvent, it is limited to the relationship between the H^+ ion and the OH^- ion. What would be useful is a more general definition that would be more applicable to other chemical reactions and, importantly, independent of H_2O .

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, a Brønsted-Lowry acid is a proton donor (PD), while a Brønsted-Lowry base is a proton acceptor (PA).

The Brønsted-Lowry definition covers the Arrhenius definition of acids and bases. Consider the prototypical Arrhenius acid-base reaction:

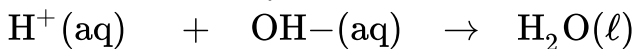


The acid species and base species are marked. The proton, however, is (by definition) a proton donor (labelled PD), while the OH^- ion is acting as the proton acceptor (labelled PA):



(PD) (PA)

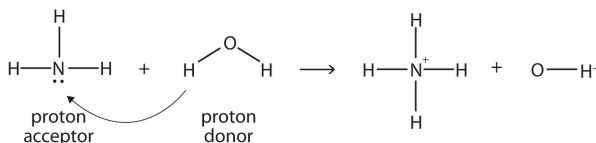
The proton donor is a Brønsted-Lowry acid, and the proton acceptor is the Brønsted-Lowry base:



(BL acid) (BL base)

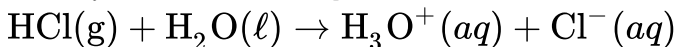
Thus H^+ is an acid by both definitions, and OH^- is a base by both definitions.

Ammonia (NH_3) is a base even though it does not contain OH^- ions in its formula. Instead, it generates OH^- ions as the product of a proton-transfer-reaction with H_2O molecules; NH_3 acts like a Brønsted-Lowry base, and H_2O acts like a Brønsted-Lowry acid:



A reaction with water is called **hydrolysis**; we say that NH_3 hydrolyses to make NH_4^+ ions and OH^- ions, as shown above.

Even the dissolving of an Arrhenius acid in water can be considered a Brønsted-Lowry acid-base reaction. Consider the process of dissolving $\text{HCl}(\text{g})$ in water to make an aqueous solution of hydrochloric acid. The process can be written as follows:



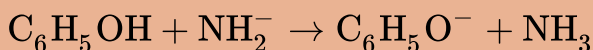
$\text{HCl}(\text{g})$ is the proton donor and, therefore, a Brønsted-Lowry acid, while H_2O is the proton acceptor and a Brønsted-Lowry base. These two examples show that H_2O can act as both a proton

donor and a proton acceptor, depending on what other substance is in the chemical reaction. A substance that can act as a proton donor or a proton acceptor is called **amphiprotic**. Water is probably the most common amphiprotic substance we will encounter, but other substances are also amphiprotic.

Examples 6.1.2

Problem

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.



Solution

The $\text{C}_6\text{H}_5\text{OH}$ molecule is losing an H^+ ; it is the proton donor and the Brønsted-Lowry acid. The NH_2^- ion (called the amide ion) is accepting the H^+ ion to become NH_3 , so it is the Brønsted-Lowry base.

Test Yourself

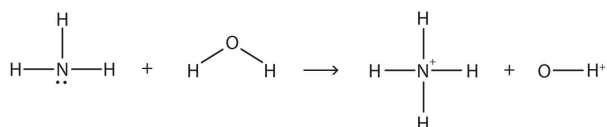
Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.



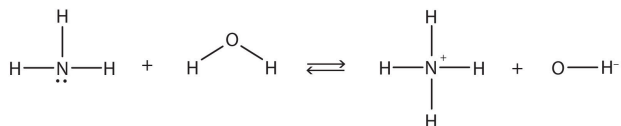
Answer

Brønsted-Lowry acid: $\text{Al}(\text{H}_2\text{O})_6^{3+}$; Brønsted-Lowry base: H_2O

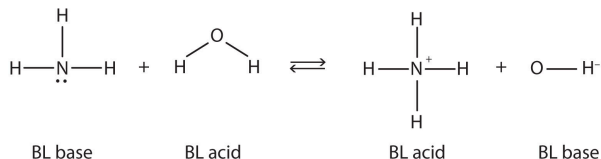
Observe the reaction between NH_3 and H_2O :



The chemical reaction does not go to completion; rather, the reverse process occurs as well, and eventually, the two processes cancel out any additional change. At this point, we say the chemical reaction is at *equilibrium*. Both processes still occur, but any net change by one process is countered by the same net change by the other process; it is a *dynamic*, rather than a *static*, equilibrium. Because both reactions are occurring, it makes sense to use a double arrow instead of a single arrow:

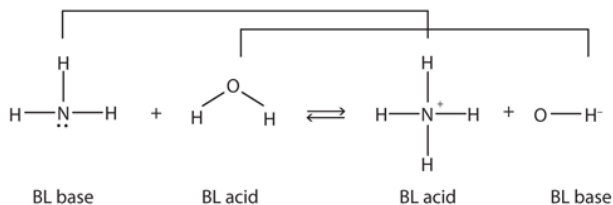


What do you notice about the reverse reaction? The NH_4^+ ion is donating a proton to the OH^- ion, which is accepting it. This means that the NH_4^+ ion is acting as the proton donor, or Brønsted-Lowry acid, while OH^- ion, the proton acceptor, is acting as a Brønsted-Lowry base. The reverse reaction is also a Brønsted-Lowry acid-base reaction:



This means that both reactions are acid-base reactions by the Brønsted-Lowry definition. If you consider the species in this chemical reaction, two sets of similar species exist on both sides. Within each set, the two species differ by a proton in their formulas, and one member of the set is a Brønsted-Lowry acid, while the

other member is a Brønsted-Lowry base. These sets are marked here:

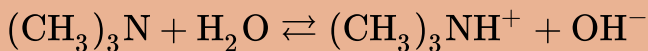


The two sets – $\text{NH}_3/\text{NH}_4^+$ and $\text{H}_2\text{O}/\text{OH}^-$ – are called **conjugate acid-base pairs**. We say that NH_4^+ is the conjugate acid of NH_3 , OH^- is the conjugate base of H_2O , and so forth. Every Brønsted-Lowry acid-base reaction can be labelled with two conjugate acid-base pairs.

Examples 6.1.3

Problem

Identify the conjugate acid-base pairs in this equilibrium.



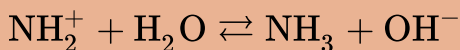
Solution

One pair is H_2O and OH^- , where H_2O has one

more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base. The other pair consists of $(\text{CH}_3)_3\text{N}$ and $(\text{CH}_3)_3\text{NH}^+$, where $(\text{CH}_3)_3\text{NH}^+$ is the conjugate acid (it has an additional proton) and $(\text{CH}_3)_3\text{N}$ is the conjugate base.

Test Yourself

Identify the conjugate acid-base pairs in this equilibrium.



Answer

H_2O (acid) and OH^- (base); NH_2^+ (base) and NH_3 (acid)

Household Acids and Bases

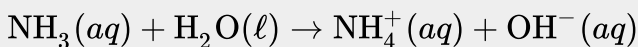
Many household products are acids or bases. For example, the owner of a swimming pool may use muriatic acid to clean the pool. Muriatic acid is another name for $\text{HCl}(aq)$. Vinegar is a dilute solution of

acetic acid $\text{HC}_2\text{H}_3\text{O}_2(aq)$ In a medicine chest, one may find a bottle of vitamin C tablets; the chemical name of vitamin C is ascorbic acid ($\text{HC}_6\text{H}_7\text{O}_6$).



Figure 6.1.1 Household acids and bases: Bowl of baking soda, vinegar, and lemon. Image attribution: Bowl of baking soda, vinegar, cleaning brush, sponge and lemon on white background © Pixel-Shot-stock.adobe.com.

One of the more familiar household bases is NH_3 , which is found in numerous cleaning products. NH_3 is a base because it increases the OH^- ion concentration by reacting with H_2O :



Many soaps are also slightly basic because they

contain compounds that act as Brønsted-Lowry bases, accepting protons from H_2O and forming excess OH^- ions. This is one explanation for why soap solutions are slippery.

Perhaps the most dangerous household chemical is the lye-based drain cleaner. Lye is a common name for NaOH , although it is also used as a synonym for KOH . Lye is an extremely caustic chemical that can react with grease, hair, food particles, and other substances that may build up and clog a water pipe.

Watch the following simulation about acids and bases.



One or more interactive elements has been excluded from this version of the text. You can view them online

here:

[https://rmit.pressbooks.pub/](https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=314#iframe-phet-1)

[rmitchemistrybridgingcourse/?p=314#iframe-phet-1](https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=314#iframe-phet-1)

Key Takeaways

- An Arrhenius acid is a compound that increases the H^+ ion concentration in aqueous solution.

- An Arrhenius base is a compound that increases the OH^- ion concentration in aqueous solution.
- The reaction between an Arrhenius acid and an Arrhenius base is called neutralisation and results in the formation of water and a salt.
- A Brønsted-Lowry acid is a proton donor; a Brønsted-Lowry base is a proton acceptor.
- Acid-base reactions include two sets of conjugate acid-base pairs.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

Media Attributions

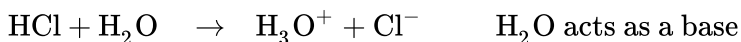
- Bowl of baking soda, vinegar, cleaning brush, sponge and lemon on white background © Pixel-Shot-stock.adobe.com

6.2 Autoionisation of Water

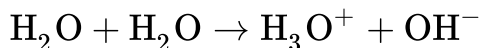
Learning Objectives

- Describe the autoionisation of water.
- Calculate the concentrations of H^+ and OH^- in solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base, as shown below:



It may not surprise you to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:



This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionisation of water**. At this level, the concentration of both $\text{H}^+ (aq)$ and $\text{OH}^- (aq)$ in a sample of pure H_2O is about 1.0×10^{-7} M. If we use square brackets around a dissolved species to imply the molar concentration of that species, we have:

$$[\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

for any sample of pure water because H_2O can act as both

an acid and a base. The product of these two concentrations is 1.0×10^{-14} as shown in the following equation:

$$[\text{H}^+] \times [\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

In acids, the concentration of $\text{H}^+ (aq)$ – written as $[\text{H}^+]$ – is greater than 1.0×10^{-7} M, while for bases the concentration of $\text{OH}^- (aq)$ – $[\text{OH}^-]$ – is greater than 1.0×10^{-7} M. However, the *product* of the two concentrations – $[\text{H}^+][\text{OH}^-]$ – is *always* equal to 1.0×10^{-14} , no matter whether the aqueous solution is an acid, a base, or neutral, which can be seen from the following:

$$[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionisation constant of water** and is denoted K_w as shown by the following equation:

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

This means that if you know $[\text{H}^+]$ for a solution, you can calculate what $[\text{OH}^-]$ has to be for the product to equal 1.0×10^{-14} , or if you know $[\text{OH}^-]$, you can calculate $[\text{H}^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w .

Example 6.2.1

Problem

What is $[\text{OH}^-]$ of an aqueous solution if $[\text{H}^+]$ is $1.0 \times 10^{-4} \text{ M}$?

Solution

Using the expression and known value for K_w :

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} = (1.0 \times 10^{-4})[\text{OH}^-]$$

We solve by dividing both sides of the equation by 1.0×10^{-4} :

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} \text{ M}$$

It is assumed that the concentration unit is molarity, so $[\text{OH}^-]$ is $1.0 \times 10^{-10} \text{ M}$.

Test Yourself

What is $[\text{H}^+]$ of an aqueous solution if $[\text{OH}^-]$ is $1.0 \times 10^{-9} \text{ M}$?

Answer

$$1.0 \times 10^{-5} \text{ M}$$

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H^+ or OH^- ions in the formula unit because $[\text{H}^+]$ or $[\text{OH}^-]$ may not be the same as the concentration of the acid or base itself.

Example 6.2.2

Problem

What is $[\text{H}^+]$ in a 0.0044 M solution of $\text{Ca}(\text{OH})_2$?

Solution

We begin by determining $[\text{OH}^-]$.

The concentration of the solute is 0.0044 M, but because $\text{Ca}(\text{OH})_2$ is a strong base, there are two OH^- ions in solution for every formula unit dissolved, so the actual $[\text{OH}^-]$ is two times this, or $2 \times 0.0044 \text{ M} = 0.0088 \text{ M}$.

Now we can use the K_w expression:

$$[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} = [\text{H}^+](0.0088 \text{ M})$$

Dividing both sides by 0.0088:

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{0.0088} = 1.1 \times 10^{-12} \text{ M}$$

$[\text{H}^+]$ has decreased significantly in this basic solution.

Test Yourself

What is $[\text{OH}^-]$ in a 0.00032 M solution of H_2SO_4 ?
(Hint: assume both H^+ ions ionise.)

Answer

$$1.6 \times 10^{-11} \text{ M}$$

For strong acids and bases, $[\text{H}^+]$ and $[\text{OH}^-]$ can be determined directly from the concentration of the acid or base itself because these ions are 100% ionised by definition. However, for weak acids and bases, this is not so. The degree, or percentage, of ionisation would need to be known before we can determine $[\text{H}^+]$ and $[\text{OH}^-]$.

Example 6.2.3

Problem

A 0.0788 M solution of $\text{HC}_2\text{H}_3\text{O}_2$ is 3.0% ionised into H^+ ions and $\text{C}_2\text{H}_3\text{O}_2^-$ ions. What is $[\text{H}^+]$ and $[\text{OH}^-]$ for this solution?

Solution

Because the acid is only 3.0% ionised, we can determine $[\text{H}^+]$ from the concentration of the acid. Recall that 3.0% is 0.030 in decimal form:

$$[\text{H}^+] = 0.030 \times 0.0788 = 0.00236 \text{ M}$$

With this $[\text{H}^+]$, then $[\text{OH}^-]$ can be calculated as follows:

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{0.00236} = 4.2 \times 10^{-12} \text{ M}$$

This is about 30 times higher than would be expected for a strong acid of the same concentration.

Test Yourself

A 0.0222 M solution of pyridine ($\text{C}_5\text{H}_5\text{N}$) is 0.44% ionised into pyridinium ions ($\text{C}_5\text{H}_5\text{NH}^+$) and OH^- ions. What are $[\text{OH}^-]$ and $[\text{H}^+]$ for this solution?

Answer

$$[\text{OH}^-] = 9.77 \times 10^{-5} \text{ M}; [\text{H}^+] = 1.02 \times 10^{-10} \text{ M}$$

Key Takeaways

- In any aqueous solution, the product of $[\text{H}^+]$ and $[\text{OH}^-]$ equals 1.0×10^{-14} .

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

6.3 The pH Scale

Learning Objectives

- Define pH.
- Determine the pH of acidic and basic solutions.

As we have seen, $[\text{H}^+]$ and $[\text{OH}^-]$ values can be markedly different from one aqueous solution to another. So, chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is a logarithmic function of $[\text{H}^+]$:

$$\text{pH} = -\log[\text{H}^+]$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[\text{H}^+]$, we can summarise as follows:

- **If $\text{pH} < 7$, then the solution is acidic.**
- **If $\text{pH} = 7$, then the solution is neutral.**
- **If $\text{pH} > 7$, then the solution is basic.**

This is known as the **pH scale**. You can use pH to make a quick determination of whether a given aqueous solution is acidic, basic, or neutral.

Example 6.3.1

Problem

Label each solution as acidic, basic, or neutral based only on the stated pH.

1. milk of magnesia, $pH = 10.5$
2. pure water, $pH = 7$
3. wine, $pH = 3.0$

Solution

1. With a pH greater than 7, milk of magnesia is basic.
(Milk of magnesia is largely $Mg(OH)_2$.)
2. Pure water, with a pH of 7, is neutral.
3. With a pH of less than 7, wine is acidic.

Test Yourself

Identify each substance as acidic, basic, or neutral based only on the stated pH.

1. human blood, $pH = 7.4$
2. household ammonia, $pH = 11.0$
3. cherries, $pH = 3.6$

Answers

1. basic
2. basic
3. acidic

Table 6.3.1 “Typical pH Values of Various Substances” gives the typical pH values of some common substances. Note that several food items are on the list, and most of them are acidic.

Table 6.3.1 Typical pH Values of Various Substances¹

Substance	pH
stomach acid	1.7
lemon juice	2.2
vinegar	2.9
soda	3.0
wine	3.5
coffee, black	5.0
milk	6.9
pure water	7.0
blood	7.4
seawater	8.5
milk of magnesia	10.5
ammonia solution	12.5
1.0 M NaOH	14.0

pH is a **logarithmic scale**. A solution that has a pH of 1.0 has 10 times the $[\text{H}^+]$ as a solution with a pH of 2.0, which in turn has 10 times the $[\text{H}^+]$ as a solution with a pH of 3.0 and so forth.

Using the definition of pH, it is also possible to calculate $[\text{H}^+]$ (and $[\text{OH}^-]$) from pH and vice versa. The general formula for determining $[\text{H}^+]$ from pH is as follows:

$$[\text{H}^+] = 10^{-\text{pH}}$$

During your studies, it will be necessary to evaluate the above expression on your calculator. This will not be covered in this book; watch this video to learn how to perform pH calculations using a calculator.

The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power of 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer, as you can see below:

$$\begin{array}{c} \text{X.YYY} \\ \swarrow \quad \searrow \\ \text{Y.YY} \times 10^x \end{array}$$

Example 6.3.2

1. Actual values may vary depending on conditions.

Problem

What are $[\text{H}^+]$ and $[\text{OH}^-]$ for an aqueous solution whose pH is 4.88?

Solution

We need to evaluate the following expression:

$$[\text{H}^+] = 10^{-4.88}$$

Depending on the calculator you use, the method for solving this problem will vary. In some cases, the “-4.88” is entered and a “10^x” key is pressed; for other calculators, the sequence of keystrokes is reversed. In any case, the correct numerical answer is as follows:

$$[\text{H}^+] = 1.3 \times 10^{-5} \text{ M}$$

Because 4.88 has two digits after the decimal point, $[\text{H}^+]$ is limited to two significant figures. From this, $[\text{OH}^-]$ can be determined:

$$[\text{OH}^-] = \frac{1 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.7 \times 10^{-10} \text{ M}$$

Test Yourself

What are $[\text{H}^+]$ and $[\text{OH}^-]$ for an aqueous solution whose pH is 10.36?

Answer

$$[\text{H}^+] = 4.4 \times 10^{-11} \text{ M}; [\text{OH}^-] = 2.3 \times 10^{-4} \text{ M}$$

There is an easier way to relate $[\text{H}^+]$ and $[\text{OH}^-]$. We can also define **pOH** in a similar way to pH:

$$\text{pOH} = -\log[\text{OH}^-]$$

In fact, p“anything” is defined as the negative logarithm of that anything. This also implies that:

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

A simple and useful relationship is that for any aqueous solution:

$$\text{pH} + \text{pOH} = 14$$

This relationship makes it straightforward to determine pH from pOH or pOH from pH and then calculate the resulting ion concentration.

Example 6.3.3

Problem

The pH of a solution is 8.22. What are pOH, $[\text{H}^+]$, and $[\text{OH}^-]$?

Solution

Because the sum of pH and pOH equals 14, we have:

$$8.22 + \text{pOH} = 14$$

Subtracting 8.22 from 14, we get:

$$\text{pOH} = 5.78$$

Now, we evaluate the following two expressions:

$$[\text{H}^+] = 10^{-8.22}$$

$$[\text{OH}^-] = 10^{-5.78}$$

So:

$$[\text{H}^+] = 6.0 \times 10^{-9} \text{ M}$$

$$[\text{OH}^-] = 1.7 \times 10^{-6} \text{ M}$$

Test Yourself

The pOH of a solution is 12.04. What are pH, $[\text{H}^+]$, and $[\text{OH}^-]$?

Answer

pH = 1.96; $[\text{H}^+] = 1.1 \times 10^{-2} \text{ M}$; $[\text{OH}^-] = 9.1 \times 10^{-13} \text{ M}$

Watch the following simulation about the pH scale.



One or more interactive elements has been excluded from this version of the text. You can view them online

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Key Takeaways

- pH is a logarithmic function of $[\text{H}^+]$.

- $[\text{H}^+]$ can be calculated directly from pH.
- pOH is related to pH and can be easily calculated from pH.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

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Transcript

6.4 Strength of Acids and Bases and their Salts

Learning Objectives

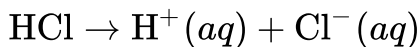
- Define a strong and a weak acid and base.
- Recognise an acid or a base as strong or weak.
- Determine if a salt produces an acidic or a basic solution.

Strong and Weak Acids and Bases

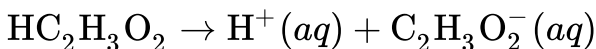
Except for their names and formulas, so far, we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way.

Consider $\text{HCl}(aq)$:

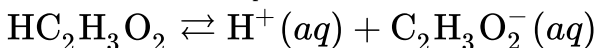
When HCl is dissolved in H_2O , it completely dissociates into $\text{H}^+(aq)$ and $\text{Cl}^-(aq)$ ions; all the HCl molecules become ions, as can be seen below:



Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. $\text{HC}_2\text{H}_3\text{O}_2$ is an example of a weak acid and its dissociation is shown below:



Because this reaction does not go 100% to completion, it is more appropriate to write it as an equilibrium, as shown below:



As it turns out, there are very few strong acids, which are given in Table 6.4.1 “Strong Acids and Bases”. If an acid is not listed here, it is a weak acid. It may be 1% ionised or 99% ionised, but it is still classified as a weak acid.

This concept applies to the bases as well: a **strong base** is a base that is 100% ionised in solution. If it is less than 100% ionised in solution, it is a **weak base**. As previously mentioned, there are not many strong bases, and in fact, any base not listed in Table 6.4.1 is considered a weak base. All strong bases are OH^- compounds. So a base based on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base.

Table 6.4.1 Strong Acids and Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	RbOH
H_2SO_4	CsOH
HClO_3	$\text{Mg}(\text{OH})_2$
HClO_4	$\text{Ca}(\text{OH})_2$
	$\text{Sr}(\text{OH})_2$
	$\text{Ba}(\text{OH})_2$

Example 6.4.1

Problem

Identify each acid or base as strong or weak.

1. HCl
2. $\text{Mg}(\text{OH})_2$
3. $\text{C}_5\text{H}_5\text{N}$

Solution

1. HCl is a strong acid as it dissociates 100% into ions.
2. $\text{Mg}(\text{OH})_2$ is a strong base as it dissociates 100% into ions.
3. The nitrogen in $\text{C}_5\text{H}_5\text{N}$ would act as a proton acceptor and, therefore, can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

Test Yourself

Identify each acid or base as strong or weak.

1. RbOH
2. HNO_2

Answers

1. strong base
2. weak acid

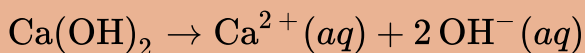
Example 6.4.2

Problem

Write the balanced chemical equation for the dissociation of strong base $\text{Ca}(\text{OH})_2$.

Solution

This is an ionic compound of Ca^{2+} ions and OH^- ions. When an ionic compound dissolves, it separates into its constituent ions:



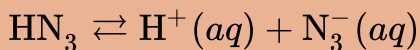
$\text{Ca}(\text{OH})_2$ is a strong base, as it dissociates 100% into ions.

Test Yourself

Write the balanced chemical equation for the dissociation of weak acid hydrazoic acid (HN_3).

Answer

The reaction is as follows:

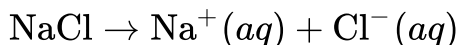


It does not proceed 100% to products because hydrazoic acid is not a strong acid.

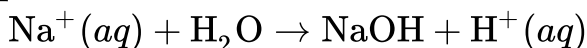
Acidity and Basicity of Salt Solutions

Certain salts will also affect the acidity or basicity of aqueous solutions because some of the ions will undergo hydrolysis, just like NH_3 does, to make a basic solution. The general rule is that salts with ions that are part of strong acids or bases will not hydrolyse, while salts with ions that are part of weak acids or bases will hydrolyse.

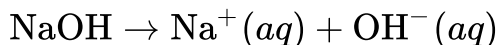
- Consider NaCl . When it dissolves in an aqueous solution, it separates into Na^+ ions and Cl^- ions, as shown below:



Will the $\text{Na}^+(aq)$ ion hydrolyse? If it does, it will interact with the OH^- ion to make NaOH :



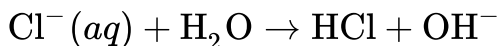
However, NaOH is a strong base, which means that it is 100% ionised in solution:



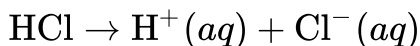
The free $\text{OH}^-(aq)$ ion reacts with the $\text{H}^+(aq)$ ion to remake a water molecule:



What is the net result? There is no change, so there is no effect on the acidity or basicity of the solution from the $\text{Na}^+(aq)$ ion. What about the Cl^- ion? Will it hydrolyse? If it does, it will take an H^+ ion from a water molecule, as shown below:



However, HCl is a strong acid, which means that it is 100% ionised in solution:

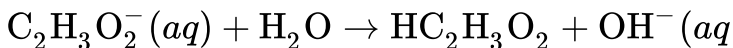


The free $\text{H}^+(aq)$ ion reacts with the $\text{OH}^-(aq)$ ion to remake a water molecule:



What is the net result? There is no change, so there is no effect on the acidity or basicity of the solution from the $\text{Cl}^-(aq)$ ion. Because neither ion in NaCl affects the acidity or basicity of the solution, NaCl is an example of a **neutral salt**.

- Let's consider a salt like $\text{NaC}_2\text{H}_3\text{O}_2$. We already know that the Na^+ ion won't affect the acidity of the solution. What about the acetate ion? If it hydrolyses, it will take an H^+ from a water molecule, as shown in the following equation:



Does this happen? Yes, it does. Why? Because $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid. Any chance a weak acid has to form, it will (the same with a weak base). As some $\text{C}_2\text{H}_3\text{O}_2^-$ ions hydrolyse with H_2O to make the molecular weak acid, OH^- ions are produced. OH^- ions make solutions basic. Thus, $\text{NaC}_2\text{H}_3\text{O}_2$ solutions are slightly basic, so such a salt is called a **basic salt**.

- There are also salts whose aqueous solutions are slightly acidic. NH_4Cl is an example. When NH_4Cl is dissolved in water, it separates into NH_4^+ ions and Cl^- ions. We have already seen that the Cl^- ion does not hydrolyse. However, the NH_4^+ ion will, as can be seen in the following equation:



The H_3O^+ ion is the hydronium ion, the more chemically proper way to represent the H^+ ion. This is the classic acid species in solution, so a solution of $\text{NH}_4^+(aq)$ ions is slightly acidic.

NH_4Cl is an example of an **acid salt**. The molecule NH_3 is a weak base, and it will form when it can, just like a weak acid will form when it can.

So there are two general rules:

1. If an ion derives from a strong acid or base, it will not affect the acidity of the solution.
2. If an ion derives from a weak acid, it will make the solution basic; if an ion derives from a weak base, it will make the solution acidic.

Example 6.4.3

Problem

Identify each salt as acidic, basic, or neutral.

1. KCl
2. KNO_2
3. NH_4Br

Solution

1. The ions from KCl derive from a strong acid (

HCl) and a strong base (KOH). Therefore, neither ion will affect the acidity of the solution, so KCl is a neutral salt.

2. Although the K^+ ion derives from a strong base (KOH), the NO_2^- ion derives from a weak acid (HNO_2). Therefore, the solution will be basic, and KNO_2 is a basic salt.
3. Although the Br^- ions derive from a strong acid (HBr), the NH_4^+ ion derives from a weak base (NH_3), so the solution will be acidic, and NH_4Br is an acidic salt.

Test Yourself

Identify each salt as acidic, basic, or neutral.

1. $(\text{C}_5\text{H}_5\text{NH})\text{Cl}$
2. Na_2SO_3

Answers

1. acidic
2. basic

Some salts are composed of ions that come from both weak acids and weak bases. The overall effect on an aqueous solution depends on which ion exerts more influence on the overall acidity. We will not consider such salts in this book.

Key Takeaways

- Strong acids and bases are 100% ionised in an aqueous solution.
- Weak acids and bases are less than 100% ionised in aqueous solutions.
- Salts of weak acids or bases can affect the acidity or basicity of their aqueous solutions.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online [here](#):

[https://rmit.pressbooks.pub/
rmitchemistrybridgingcourse/?p=324#h5p-42](https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=324#h5p-42)

Transcript

6.5 Acid-Base Titrations

Learning Objectives

- Describe a titration experiment.
- Explain what an indicator does.
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is that they are often used to determine quantitative amounts of acid or the base used for the experiment. Performing chemical reactions quantitatively to determine the exact amount of a **reagent** is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

In a titration, one reagent has a known concentration or amount, while the other reagent has an unknown concentration or amount. Typically, the known reagent (the **titrant**) is added to the unknown quantity and is dissolved in solution. The unknown amount of substance (the **analyte**) may or may not be dissolved in the solution (But usually it is). The titrant is added to the analyte using a precisely calibrated volumetric delivery tube called a burette (Figure 6.5.1 “Equipment for Titrations”). The burette has markings to determine how much volume of solution has been added to the analyte. When the reaction is complete, it is said to be at the **equivalence point**; the

number of moles of titrant can be calculated from the concentration and the volume, and the balanced chemical equation can be used to determine the number of moles (and then concentration or mass) of the unknown reactant.

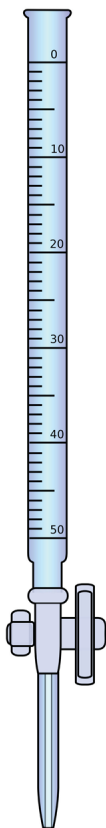


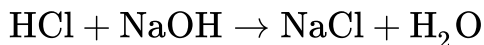
Figure 6.5.1 “Equipment for Titrations.” A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License

For example, suppose 25.66mL (or 0.02566L) of 0.1078M HCl was used to titrate an unknown sample of NaOH . What mass of NaOH was in the sample?

We can calculate the number of moles of HCl reacted:

$$\# \text{ mol HCl} = (0.02566 \text{ L})(0.1078 \text{ M}) = 0.002766 \text{ mol HCl}$$

We know the balanced chemical reaction between HCl and NaOH :



So, we can construct a conversion factor to convert to the number of moles of NaOH reacted:

$$0.002766 \cancel{\text{ mol HCl}} \times \frac{1 \text{ mol NaOH}}{1 \cancel{\text{ mol HCl}}} = 0.002766 \text{ mol NaOH}$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766 \cancel{\text{ mol NaOH}} \times \frac{40.00 \text{ g NaOH}}{1 \cancel{\text{ mol NaOH}}} = 0.1106 \text{ g NaOH}$$

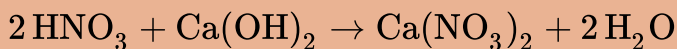
This type of calculation is performed as part of a titration.

Example 6.5.1

Problem

What mass of $\text{Ca}(\text{OH})_2$ is present in a sample if it is titrated to its equivalence point with 44.02mL of

0.0885 M HNO_3 ? The balanced chemical equation is as follows:



Solution

In litres, the volume is 0.04402 L . We calculate the number of moles of titrant:

$$\# \text{ moles HNO}_3 = (0.04402\text{ L})(0.0885\text{ M}) = 0.00390\text{ mol HNO}_3$$

Using the balanced chemical equation, we can determine the number of moles of Ca(OH)_2 present in the analyte:

$$0.00390 \cancel{\text{ mol HNO}_3} \times \frac{1 \text{ mol Ca(OH)}_2}{2 \cancel{\text{ mol HNO}_3}} = 0.00195 \text{ mol Ca(OH)}_2$$

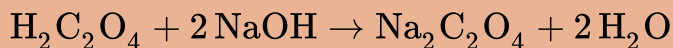
Then we convert this to a mass using the molar mass of Ca(OH)_2 :

$$0.00195 \cancel{\text{ mol Ca(OH)}_2} \times \frac{74.1 \text{ g Ca(OH)}_2}{1 \cancel{\text{ mol Ca(OH)}_2}} = 0.144 \text{ g Ca(OH)}_2$$

Test Yourself

What mass of $\text{H}_2\text{C}_2\text{O}_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of

0.2235M NaOH? The balanced chemical reaction is as follows:



Answer

0.182 g

How does one know if a reaction is at its equivalence point? Usually, the person performing the titration adds a small amount of an **indicator**, a substance that changes colour depending on the acidity or basicity of the solution. Because different indicators change colours at different levels of acidity, choosing the correct one is important in performing an accurate titration.

Key Takeaways

- A titration is the quantitative reaction of an acid and a base.
- Indicators are used to show that all the analyte has reacted with the titrant.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=327#h5p-43>

Transcript

6.6 Buffers

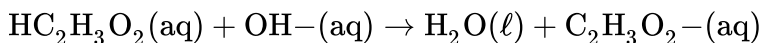
Learning Objectives

- Define a *buffer*.
- Correctly identify the two components of a buffer.

Weak acids are relatively common, even in the foods we eat. But we occasionally encounter a strong acid or base, such as stomach acid, which has a strongly acidic pH of 1.7. By definition, strong acids and bases can produce a relatively large amount of H^+ or OH^- ions and consequently have marked chemical activities. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [approximated as $0.1\text{M HCl}(aq)$] were added to the bloodstream and no correcting mechanism were present, the pH of the blood would decrease from about 7.4 to about 4.7 — a pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimising such dramatic pH changes.

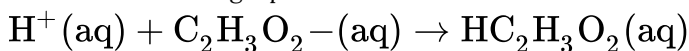
This mechanism involves a **buffer**, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved $\text{HC}_2\text{H}_3\text{O}_2$ (a weak acid) and $\text{NaC}_2\text{H}_3\text{O}_2$ (the salt derived from that weak acid). Another example of a buffer is a solution containing NH_3 (a weak base) and NH_4Cl (a salt derived from that weak base).

Let us use an $\text{HC}_2\text{H}_3\text{O}_2$ / $\text{NaC}_2\text{H}_3\text{O}_2$ buffer to demonstrate how buffers work. If a strong base – a source of OH^- (*aq*) ions – is added to the buffer solution, those OH^- ions will react with the $\text{HC}_2\text{H}_3\text{O}_2$ in an acid-base reaction, as shown below:



Rather than changing the pH dramatically by making the solution basic, the added OH^- ions react to make H_2O , so the pH does not change much.

If a strong acid – a source of H^+ ions – is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid, it is not ionised much. This means that if lots of H^+ ions and $\text{C}_2\text{H}_3\text{O}_2^-$ ions are present in the same solution, they will come together to make $\text{HC}_2\text{H}_3\text{O}_2$, as shown in the following equation:



Rather than changing the pH dramatically and making the solution acidic, the added H^+ ions react to make molecules of a weak acid. Figure 6.6.1 "The Actions of Buffers" illustrates both actions of a buffer.

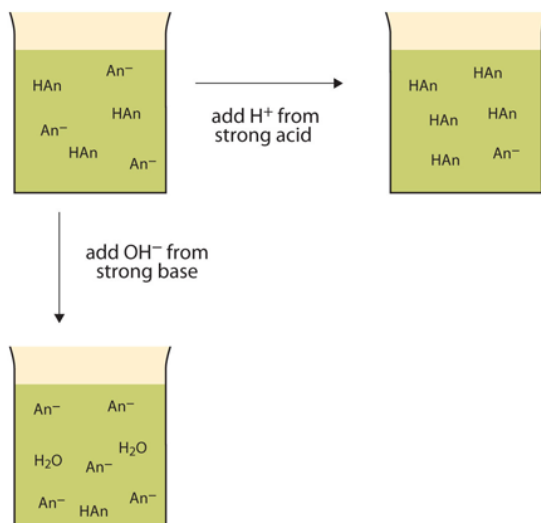
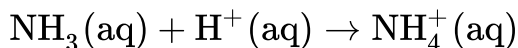
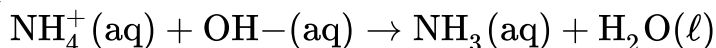


Figure 6.6.1 "The Actions of Buffers." Buffers can react with both strong acids (top) and strong bases (side) to minimise large changes in pH. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

Buffers made from weak bases and the salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl , NH_3 molecules can react with any excess H^+ ions introduced by strong acids, as shown below:



While the $\text{NH}_4^+(\text{aq})$ ion can react with any OH^- ions introduced by strong bases, as can be seen in the following equation:



Problem

Which combinations of compounds can make a buffer solution?

1. HCHO_2 and NaCHO_2
2. HCl and NaCl
3. CH_3NH_2 and $\text{CH}_3\text{NH}_3\text{Cl}$
4. NH_3 and NaOH

Solution

1. HCHO_2 is formic acid, a weak acid, while NaCHO_2 is the salt made from the anion of the weak acid (the formate ion $[\text{ce}{\text{CHO}_2}^{\text{-}}]$). The combination of these two solutes would make a buffer solution.
2. HCl is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
3. CH_3NH_2 is methylamine, which is like NH_3 ,

with one of its **H** atoms substituted with a **CH₃** group. It is a weak base. The compound **CH₃NH₃Cl** is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.

4. **NH₃** is a weak base, but **NaOH** is a strong base. The combination of these two solutes would not make a buffer solution.

Test Yourself

Which combinations of compounds can make a buffer solution?

1. **NaHCO₃** and **NaCl**
2. **H₃PO₄** and **NaH₂PO₄**
3. **NH₃** and **(NH₄)₃PO₄**
4. **NaOH** and **NaCl**

Answers

1. no
2. yes
3. yes
4. no

Buffers work well only for limited amounts of added strong acid or base. Once either solute is completely reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say

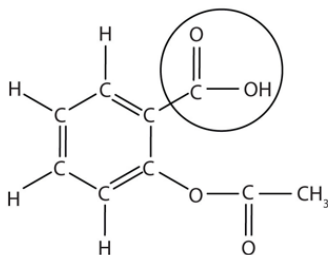
that a buffer has a certain **capacity**. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimise extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 (the second compound is another way to write $\text{CO}_2(aq)$). With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

The Acid That Eases Pain

Let's take a look at an acid that is probably the most common medicine: acetylsalicylic acid, also known as aspirin. Aspirin is well known as a pain reliever and antipyretic (fever reducer).

The structure of aspirin is shown in the accompanying figure. The acid part is circled; it is the hydrogen atom in that part that can be donated as aspirin acts as a **Brønsted-Lowry acid**. Acetylsalicylic acid is a weak acid. However, it is still an acid, and given that some people consume relatively large amounts of aspirin daily, its acidic nature can cause problems in the stomach lining despite the stomach's defences against its own stomach acid.



Because the acid properties of aspirin may be problematic, many aspirin brands offer a “buffered aspirin” form of the medicine. In these cases, the aspirin also contains a buffering agent – usually **MgO** – that regulates the acidity of the aspirin to minimise its acidic side effects.

Key Takeaways

- A buffer is a solution that resists sudden changes in pH.

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=331#h5p-44>

Transcript

PART VII

CHAPTER 7: CHEMICAL ENERGY AND RATES OF REACTIONS

7.1 Forms of Energy

Learning Objectives

- Define the law of conservation of matter and energy.
- Understand the different ways energy can be transformed into alternative forms.
- Define the study of thermochemistry and the types of energy of interest within physical chemistry.

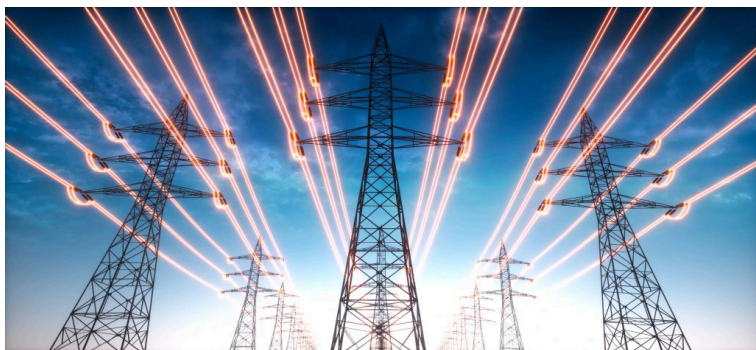


Image attribution: Electricity transmission towers with glowing wires against blue sky – Energy concept © peterschreiber.media – stock.adobe.com.

When thinking about energy, your mind may naturally go to electricity – and why shouldn't it! Electrical energy powers modern society (literally!). But where does it come from? **The law of**

conservation of matter and energy states that neither can be created nor destroyed – it can only be transformed. If that's the case, how is electrical energy generated?

Energy Transformations

Let's take the example of a coal-fired power plant. To produce energy, the power plant needs to generate an electromotive force through the spinning of steam **turbines**. These turbines rotate through the use of pressurised steam, which is produced by raising the temperature of water until it boils. This water needs heat in order to undergo a change in its state of matter – supplied by burning coal through combustion (see Figure 7.1.1).

Coal Power Electrical Station

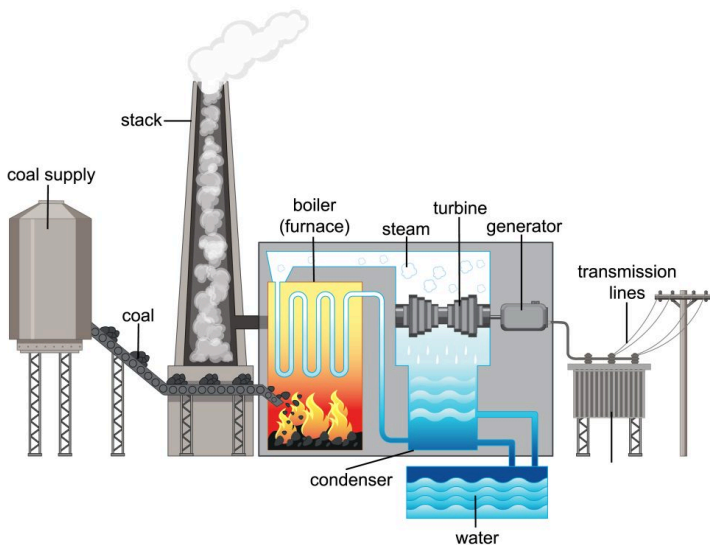


Figure 7.1.1: Diagram showing a typical process of a coal power plant.
Transcript. Image attribution: Diagram showing Coal Power Electrical Station © blueringmedia – stock.adobe.com.

Consider the energy transformations that occur during this process: electrical energy is generated through the mechanical energy of the steam turbine, which is energy provided by the energy of the pressurised steam. The steam was generated by supplying thermal energy to the water to allow the molecules to overcome **intermolecular forces** and form a gas. The thermal energy was provided by harnessing the coal's potential chemical energy through a reaction. Energy transformations are continuously occurring in our everyday lives, from cooking and eating to the nuclear fusion that occurs in the sun.

We can even take this one step further. The coal found in the Earth's crust was generated through the intense heat and compression of carbon-based lifeforms, such as plants. For millions of years, those plants received light energy from the sun. Coal originally gets some of its energy from the sun! (Although it might not be renewable like others).

Energy transformations are all around us! Have a look at this interactive simulation to see the different types of energy transfers that can happen in daily processes. (Be sure to tick the 'energy symbols' box before you start.)



One or more interactive elements has been excluded from this version of the text. You can view them online

here: <https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=629>

Simulation by PhET Interactive Simulations, University of Colorado

Forms of Energy

So, what exactly is energy? Energy is the ability to do **work** and is typically located within matter ¹. Work is performed against a force, such as gravity or electrical resistance. Here are some forms which energy can take:

1. ; photons are a notable exemption

Kinetic Energy	Energy found within movement of particles and objects.	A car moving at speed.
Gravitational Potential Energy	Potential energy within objects that are positioned against a gravitational field.	A ball held upwards into the sky.
Chemical Energy	Energy harnessed through the rearrangement of atoms and molecules.	Covalent and Ionic Bonds
Light Energy	Energy found within photons and electromagnetic waves.	Infrared radiation from the sun.
Sound Energy	Energy associated with sound and air pressure.	The soundwaves produced from a speaker.
Thermal Energy	The heat energy found within hot objects.	The sensation of heat from holding a hot cup.
Nuclear Energy	Energy harnessed through the rearrangement of protons and neutrons.	The nuclear fission of uranium.

In chemistry, we are mainly concerned with the interconversion of chemical energy and thermal energy. This field is specifically known as



Figure 7.1.2: Metallic sodium reacts quite violently with a rather tame substance: water. The reaction often causes a small explosion! Image attribution: Metallic sodium in water © Alexandre – stock.adobe.com i

[pb_glossaryid="663"]thermochemistry[/pb_glossary].

The available energy from seemingly unreactive sources is tremendous when considering all of the subatomic particles that compose matter (see Figure 7.1.2). The ways in which we harness that energy are dependent on the substances. We've already seen how combustion reactions can be utilised, but a variety of other reactions, such as **redox**, neutralisation and a wide variety of organic substitutions and reactions can also produce thermal energy.

As you will soon discover, heat plays an essential role in determining the extent and rates of reactions. It also is a major factor in chemical decomposition and other physical properties. Chemicals with a low **flashpoint** may ignite unexpectedly in experiments that involve temperature increase! As such, the management of energy is essential for any aspiring chemist.

Key Takeaways

- According to the law of conservation of energy and matter; energy can only be transformed, never created.
- Common forms of energy include kinetic, gravitational, chemical, light, sound, thermal and nuclear.
- Thermochemistry studies the interconversion of chemical and thermal energy.

Practice Questions



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Transcript

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- Metallic sodium in water © Alexandre - stock.adobe.com is licensed under a All Rights Reserved license

7.2 Thermochemistry Essentials

Learning Objectives

- Define how temperature and heat are understood on the molecular level by Collision Theory.
- Learn about the absolute scale of the Kelvin and its relationship to degrees Celsius.
- Define the three types of heat transfer: conduction, convection and radiation.
- Define the three types of thermodynamic systems: open, closed and isolated, and the allowed transfer of matter and energy in each.

Collision Theory

To understand thermochemistry, we must first talk about the basics of **collision theory**. Collision theory is a chemical model which applies physics principles to understand the mechanisms of reactions. Let's consider the formation of water: $2\text{H}_2 + \text{O}_2 \longrightarrow 2\text{H}_2\text{O}$. Under this model, we will represent all molecules as balls, as demonstrated in Figure 7.2.1.

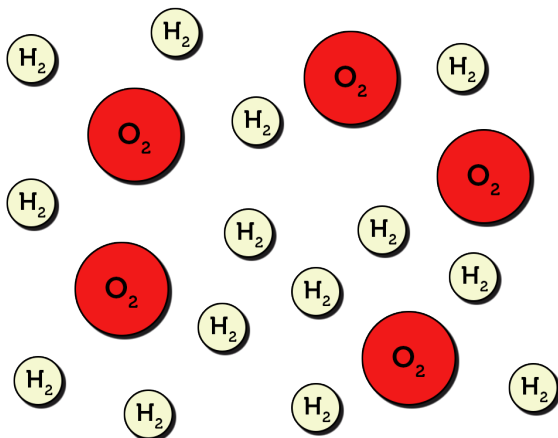


Figure 7.2.1: Oxygen and hydrogen molecules floating in space.

From your own experience (in the lab or cooking on a stove), you may know that the temperature of a **system** correlates to the rate of a reaction. According to collision theory, **temperature** is directly related to the **kinetic energy** of molecules. As temperature rises, the speed at which molecules move increases (see Figure 7.2.2).

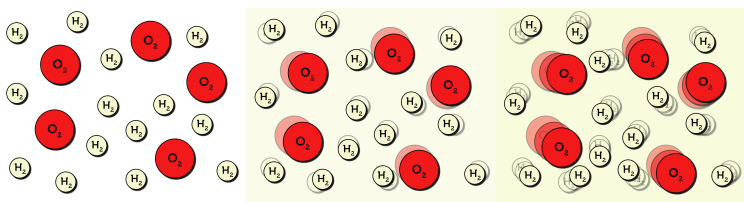


Figure 7.2.2: As substances heat up, molecules move faster and faster.

With molecules moving faster, they are more likely to collide with each other with enough energy, causing a reaction to take place. Higher temperatures allow more successful collisions to occur more often, increasing the reaction rate.

The Kelvin

With this understanding, we must revisit how temperature is traditionally measured. The majority of the world utilises degrees **Celsius** ($^{\circ}\text{C}$) – a system based on the freezing (0°C) and boiling (100°C) point of water. However, a reading of 0°C does not mean that no energy exists within a system; a lower temperature, such as -5°C , can still be achieved. For our energy calculations, an absolute value is needed: the **Kelvin** (K). The conversion from degrees Celsius to Kelvin is as follows:

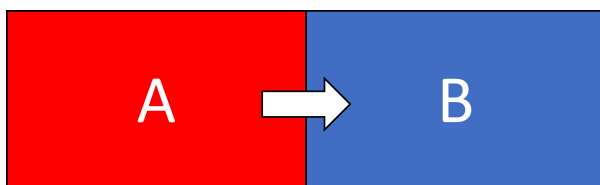
$$K = ^{\circ}\text{C} + 273.15$$

Under the Kelvin system, a value of 0 K represents that a molecule has no kinetic energy and is completely still. This is known as **absolute zero**¹. Energy calculations within science are generally performed in Kelvin for this reason.

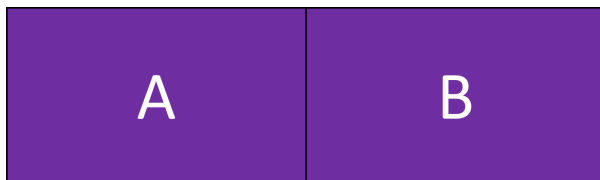
Conduction

With our understanding of temperature being a measure of kinetic energy, we can appreciate how **heat** occurs. Heat is the transfer of energy from one body to another. This can only occur when a temperature difference is present. A hotter body will impart its energy onto a colder body until **thermodynamic equilibrium** is established. As heat is transferred, the hotter body will cool down as it loses its kinetic energy, while the colder body will begin to rise in temperature (see Figure 7.2.3).

1. . Absolute zero is a theoretical concept and cannot physically be achieved.



Transfer of thermal energy



Equilibrium

Figure 7.2.3: All bodies will change temperature until they reach thermodynamic equilibrium – where both contain the same amount of thermal energy.



Figure 7.2.4: Holding an ice cube feels cold, as the thermal energy within your hand is being transferred to the ice cube. That thermal energy is enough to melt the ice. Image attribution: child holding ice cubes – winter © BarbaraKrupa – stock.adobe.com.

Therefore, what we consider to be hot or cold isn't due to the temperature itself – but the temperature difference. Something appears to feel cold because our hands are hotter than it is (see Figure 7.2.4), and vice versa. Our morning coffee on a particularly chilly day might appear to be warmer than usual because our bodies are colder than normal. This is also why some people find it hard to check their own

temperature by touch when feverish – if their whole body is getting

hotter, nothing will feel out of the ordinary even though body temperatures have risen.

What we have just described is **conduction**: the movement of thermal energy between two bodies in physical contact. It is the simplest of three main ways to transfer energy. From cooking a steak on a cast-iron pan to the cold feeling of placing an icepack on your body – conduction is the most direct form of heat transfer.

Convection

Thermal transfer is, however, not limited to solid objects. When a pot of water is heated up, how does that thermal energy spread itself around the container?

All **fluids** exhibit convection. As liquids or gasses are heated, hotter areas become less dense and begin to rise, allowing colder molecules to replace them and be heated themselves. This movement creates a **convection current** (see Figure 7.2.5).

An air-fryer or convection oven uses this concept to its advantage. By heating up food through a convection current of air, a more uniform temperature can be achieved. While a frying pan can only heat food from one side at a time, a convection oven can heat from everywhere at once. To assist in this movement, fans are used to improve efficiency – allowing cooking temperatures to be reduced in fan-forced systems.

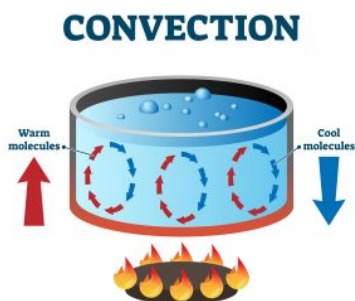


Figure 7.2.5: Molecules warmed through convection become less dense and rise to the top of the system, where colder molecules take their place. This flow is known as a convection current. Image attribution: Convection currents vector illustration labeled diagram © VectorMine – stock.adobe.com

Radiation

So far, the models we have looked at rely on matter to be present to transfer thermal energy. However, there is no matter in space, so how does the sun transfer heat to our planet?

The sun produces a wide range of electromagnetic radiation. Our eyes are able to detect only a small range of wavelengths (known as the visible spectrum; see Figure 7.2.5) – but a variety of other wavelengths are emitted.

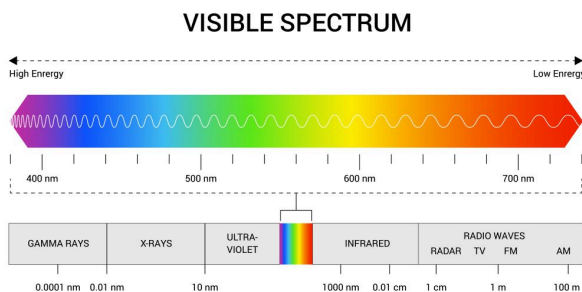


Figure 7.2.5:
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Infrared is the most common form of radiation generating heat. All objects with heat emit radiation, although hotter bodies manifest higher energy waves (see Figure 7.2.6). These wavelengths collide with molecules – causing some kinetic energy to be imparted. With more kinetic energy, a higher temperature is achieved and the object is heated.

As such, **thermal radiation** requires no physical matter between the source and a target of interest to transfer energy. In fact, all forms of electromagnetic radiation do not require a medium to travel through, explaining why the sun can impart light and energy onto the Earth even through the vacuum of space.

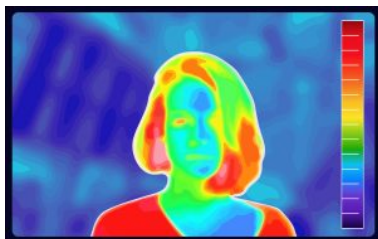


Figure 7.2.6: IR cameras are able to measure the infrared radiation emitted from hot objects. This allows for temperature to be measured from a distance and is the main technology by night-vision cameras. Image attribution: Vector graphic of Thermographic image of a woman face showing different temperatures in a range of colors from blue showing cold to red showing hot. Medical thermal imaging of human female face. © Cipta – stock.adobe.com.

HEAT TRANSFER

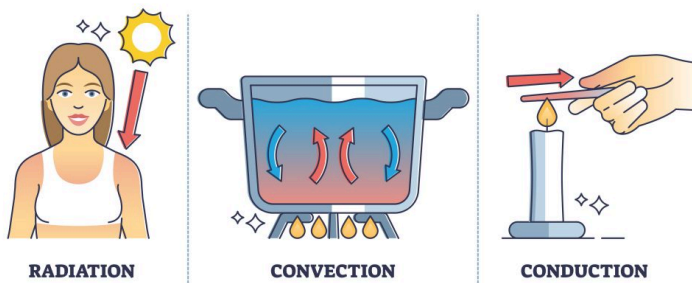


Figure 7.2.7: The 3 types of heat transfer: radiation, convection, and conduction. Image attribution: Heat transfer types with radiation, convection and conduction types outline diagram. Labeled educational scheme with thermal energy exchange methods vector illustration. Hot temperature sources list. © VectorMine – stock.adobe.com.

Systems of Heat

With all 3 forms of heat transfer discussed (see Figure 7.2.7), we can understand how different thermal systems operate. In this context, a **system** refers to the part of the **universe** under study. Between the universe and the system is its direct **surroundings** – which can exchange heat and matter depending on the system (see Figure 7.2.8). The type of exchange depends on the

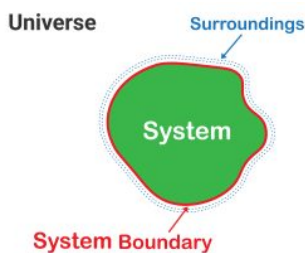


Figure 7.2.8: All thermodynamic studies have spatial positions that can be labelled. Image attribution: Thermodynamic system, boundary, system and surroundings © Reuel Sa – stock.adobe.com

system utilised. We observe three main types: open, closed, and isolated (see Figure 7.2.9).

An **open system** is one where both matter and energy can be exchanged with the environment. A boiling pot of water can be considered an open system: heat can radiate and conduct out to the surrounding environment, and water (in the form of steam) can escape. This is an important system for distillation and extraction methods. A **closed system** limits the transfer of matter but allows the transmittance of heat. By placing a lid on our pot, we are preventing matter from escaping the system as heat. We can, however, still heat up or cool down our system. An **isolated system** prevents both mass and heat from escaping into the surroundings. Heat transfer via convection can be limited through the implementation of a vacuum surrounding the system, while a reflective surface limits energy loss through radiation.

THERMOCHEMISTRY ◇◇

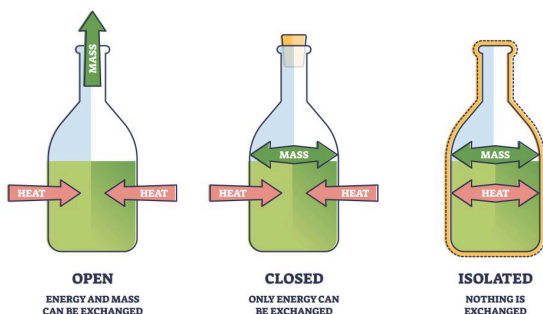


Figure 7.2.9:
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Energy studies performed within thermochemistry can involve all of these different types of systems. Understanding what flows from a system to its surroundings and the larger universe, alongside accounting for it, allows us to perform many different thermochemical calculations and experiments.

Key Takeaways


- Temperature is a measure of the kinetic energy of molecules. As temperature raises, the speed which molecules move does as well.
- In science, the Kelvin is the preferred measure of temperature when performing calculations. 0 K is known as absolute zero.
- All bodies in contact will attempt to reach thermodynamic equilibrium.
- There are three types of heat transfer: conduction (physical contact); convection (fluid currents); and radiation (electromagnetic waves).
- There are three types of systems: open (matter and energy exchange); closed (only energy exchanged), and isolated (no transfer)

Exercises

Practice Questions



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- Heat transfer types with radiation, convection and conduction

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7.3 Exothermic and Endothermic Reactions

Learning Objectives

- Define what constitutes an exothermic and endothermic chemical reaction.
- Learn about enthalpy and the expected change in exothermic and endothermic reactions.

All reactions involve the rearrangement of chemical bonds within molecules. The bonds that form as a result of a chemical reaction will hold either more or less chemical energy than the initial reactants. This is expressed as heat. You will most likely be familiar with **exothermic** and **endothermic** reactions throughout your day-to-day life — distinguished through how energy is released or absorbed.

Exothermic reactions are those that release energy into the environment. Remember how the **law of conservation of energy** requires that energy be maintained? In an exothermic reaction, the **chemical energy of products** is less than that of the starting **reactants**. This difference in energy is expressed as heat. This is the theory behind instant heat packs. When activating, reactants can interact and transform into end products with lower chemical energies. Heat is produced as a result, allowing the soothing of injuries and aches on one's body.

In contrast, endothermic reactions absorb energy from the

environment. When the reactants come into contact, the final product has chemical energy demands greater than the reactants – meaning it must pull in energy from its surroundings. This results in a cold sensation when touched.

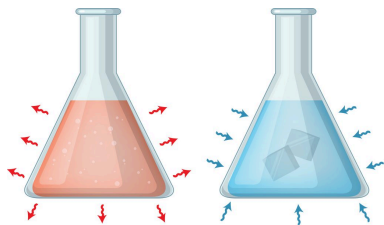
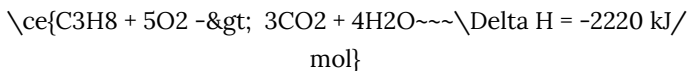


Figure 7.3.1: Exothermic and Endothermic reactions give and receive thermal energy from the environment accordingly. Image attribution: Types of chemical reactions Exothermic and endothermic reactions © Nandalal – stock.adobe.com.

Enthalpy

An important concept to chemists is the amount of energy gained or lost through chemical reactions (such as combustion). It is vital, therefore, that we define a new concept to measure this phenomenon: **enthalpy**.

Enthalpy (H) is the amount of energy found within a system (at constant pressure). Within the context of thermochemistry, we are concerned with how chemical energy transforms into thermal energy. Knowing that exothermic and endothermic reactions depend on the change of chemical energy between reactants and products, we can determine if a reaction is endothermic or exothermic in nature by observing the change in enthalpy. Let us take the combustion of propane for example (cited from W.M.Haynes CRC Handbook of Chemistry 97th Edu.):



The combustion of propane produces a negative enthalpy, denoted by ΔH . This means that compared to the reactants, the final products possess less energy. This energy is released in the form of heat. As such, a negative ΔH informs us that the reaction is exothermic. Conversely, a positive ΔH indicates an endothermic reaction. For this equation, every mole of propane combusted releases 2220 kilojoules of energy to the environment.

Key Takeaways

- Chemical reactions are either endothermic (draw energy) or exothermic (release energy).
- A chemical reaction is exothermic if the chemical energy of products is lower than that of the starting reactants, while an endothermic reaction occurs when the products' energy is higher.
- Enthalpy is the measure of energy found within a molecule or system. The change in enthalpy allows us to determine the amount of energy drawn or released from a reaction.

Exercises

Practice Questions



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Transcript

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7.4 Calorimetry and Heat Capacity

Learning Objectives

- Detail how calorimetry can be used to determine the amount of energy released in a chemical reaction.
- Discern the difference between constant-pressure (coffee cup) and constant-volume (bomb) calorimeters.
- Learn about specific and latent heat capacity and how they can be utilised in determining change in enthalpy.

Calorimetry

Determining the amount of energy within a reaction is an area of science known as calorimetry. Within calorimetry, a reaction is performed under known conditions. The enthalpy change will invoke an increase or decrease in the temperature of the environment (see Figure 7.4.1).

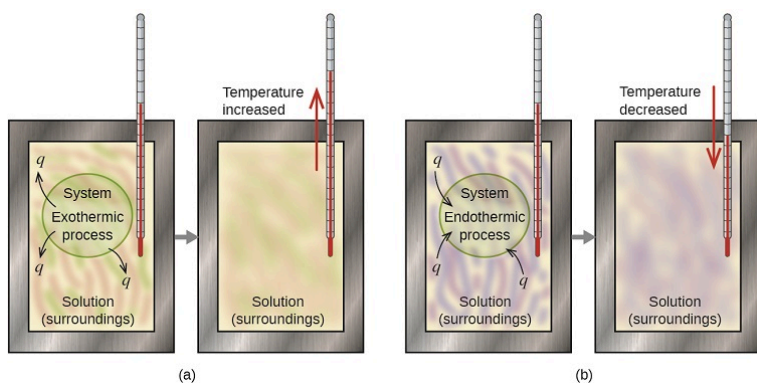


Figure 7.4.1: A calorimeter operates by measuring the energy of the environment before and after a chemical process. An exothermic reaction (a) will see a rise in the environment's temperature, while an endothermic reaction (b) will see a decrease. Image attribution: Calorimetry © Manjusha T Saraswathiamma is licensed under a CC BY-NC-SA (Attribution NonCommercial ShareAlike) .

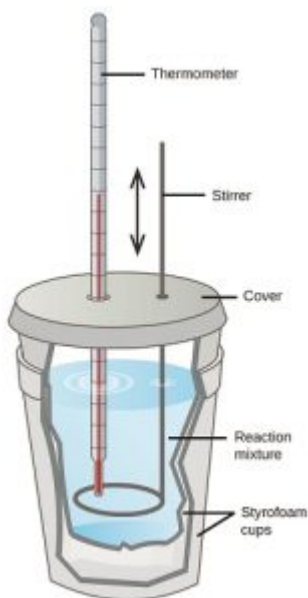


Figure 7.4.2: A coffee cup calorimeter can be easily created through stacking two polystyrene cups, a lid, stirrer and a thermometer. Image attribution: Coffee Cup Calorimeter © Manjusha T Saraswathiamma is licensed under a CC BY-NC-SA (Attribution NonCommercial ShareAlike)

There are two types of calorimeters commonly used: constant pressure and constant volume, although they are more commonly referred to by less scientific names! “Coffee cup” calorimeters (constant pressure) involve the use of an insulated container to create a closed system. This can be easily achieved in science classrooms by stacking two polystyrene cups (see Figure 7.4.2). By filling the cup with a medium (commonly water), a reaction or process can then be facilitated, with the energy released or absorbed quantifiable through a change in temperature. In the case of sodium chloride salt being dissolved in water, the **“enthalpy of dissolution”** is

positive – indicating an endothermic reaction. The temperature should, therefore, decrease. Due to the overall crudeness of this setup, energy readings will involve a degree of inaccuracy as energy seeps out from the container. Commercial options are available to alleviate these issues alongside providing more accurate electrical thermometers.

Bomb calorimeters (constant volume; see Figure 7.4.3) allow for the measurement of violent and energetic combustion reactions. A sealed “bomb” is loaded with the substance under investigation, with a source of ignition, generally an electrical spark, into an insulated container. The generated explosion will heat the water surrounding the bomb unit, allowing for energy release to be calculated. Bomb calorimeters are more accurate than their coffee cup counterparts but require calibration before use. Explosions are always exothermic processes, and so bomb calorimeters are generally useful in detecting increases in temperature.

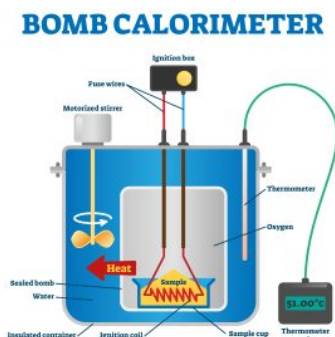


Figure 7.4.3: A bomb calorimeter allows for measurement of reactions that hold explosive potentials! Image attribution: Bomb calorimeter vector illustration. Labeled educational explain scheme. © VectorMine – stock.adobe.com.

Specific Heat Capacity

While it is easy for us to measure the temperature change of a substance, converting that figure demands that we know how much energy it takes to heat a given mass of a substance. **Specific heat capacity** is the amount of energy needed to heat 1 gram of a substance by one degree Celsius or Kelvin¹. It is used in the

1. Kelvin and Celsius can be used interchangeably, as we are

following equation:\[

$$q=mc\Delta T$$

\]

Where, q is the energy exerted or emitted (J)

m is the mass of the substance (g)

c is the specific heat capacity of the substance ($\text{JK}^{-1}\text{g}^{-1}$) or ($\text{J}^{\circ}\text{C}^{-1}\text{g}^{-1}$)

ΔT is the temperature change of the substance (K) or ($^{\circ}\text{C}$)

While different substances will have different heat capacities – commonly utilised environmental mediums have been well-documented in scientific literature. Water is the most common substance, thanks to its availability and relatively high specific heat capacity (owing to its strong hydrogen bonding). **It is recommended to learn the heat capacity of water by memory.** For any other substances, capacities will be given within the question.

Table 7.4.1: Specific

Substance
<u>water</u>
steam
ethanol
air
iron
gold
helium

measuring the change in temperature, not the absolute value. Both increase at the same rate

You may be wondering as to how these units were initially discovered. After all, to determine how much energy is needed to heat a substance, we would have needed to know how much energy was being provided initially. So, what came first?

As you may have noticed, the word ‘calorimetry’ derives from the energy unit ‘calorie’. Within the early days of this science, 1 calorie was understood to be the energy needed to heat 1 gram of water by 1 degree Celsius. From this, the heat capacities of other substances were determined. Since then, the global scientific community has pushed for the standardisation of units and measurements across fields. As such, outside of medical health and nutritional studies, scientists instead utilise joules (J) as the preferred energy unit. To convert, utilise the following relationship:

$$1 \text{ cal} = 4.184 \text{ J}$$

This book will continue to utilise the joule when referring to energy. *We can all agree however ‘Joulrimetry’ definitely doesn’t run off the tongue as well!*

Example: A calorimeter filled with 500g of water is heated from 298K to 344K. How much energy has the system absorbed?

From the question, we can determine the following quantities:

$$m = 500g$$

$$\Delta T = 344 - 298 = 46K$$

From literature, we know that the specific heat capacity of water c will be $4.184JK^{-1}g^{-1}$. We can now utilise these values in the following equation:

$$q = mc\Delta T$$

$$q = (500g) \times (4.184JK^{-1}g^{-1}) \times (46K)$$

$$q = 96232J \text{ (or } 96.232kJ)$$

This process saw 96.232 kJ of energy absorbed by the system.

Latent Heat Capacity

Whereas specific heat details the amount of energy needed to heat a substance, latent heat capacity quantifies the amount of energy needed to induce a state change within a substance. It utilises the following equation:

$$q = mL$$

Where, q is the energy exerted or emitted (J)

m is the mass of the substance (g)

L is the latent heat capacity of the substance of the transformation taking place (Jg^{-1}) or (Jg^{-1})

The type of latent heat described depends on the transformation taking place:

- Latent heat of fusion – the amount of energy required or released during solid/liquid conversion.
- Latent heat of vaporisation – the amount of energy required or released during liquid/gas conversion.
- Latent heat of sublimation – the amount of energy required or released during solid/gas conversion.

Table 7.4.2

Substance
Latent heat of fusion (Solid/Liquid)
Latent heat of vaporisation (Liquid/Gas)

Figure 7.4.4 shows the change in temperature of a substance as heat is applied. Note that during a change in state, such as in the freezing or melting process, the temperature doesn't change – energy is focused into overcoming intermolecular forces holding molecules together. Additionally, the gradients between the state changes may also be different, as the different states of matter within a substance will have different specific heat capacities.

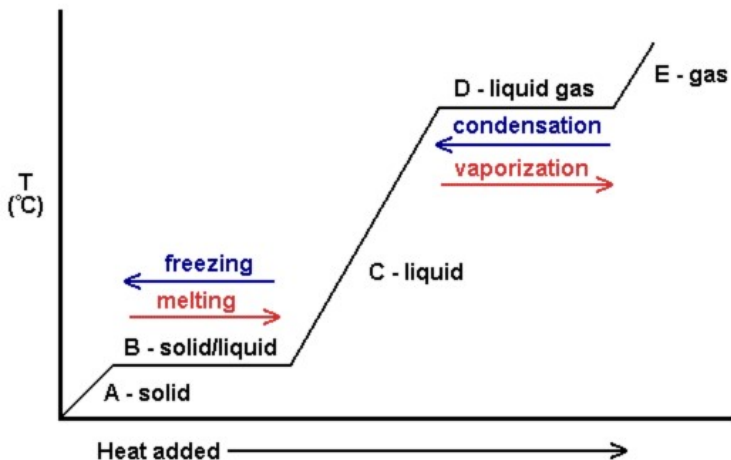


Figure 7.4.4: Example of a temperature curve. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

This principle of temperature not rising when a state transformation takes place is commonly used within chemistry for extractions and purifications but also in many other real-world applications. For instance, rice cookers use this to know when to automatically turn off. If the temperature is measured to be above 100°C , all of the water must be evaporated, as the rice and water mixture wouldn't be able to rise further if water was still present. That must mean the rice is fully cooked!

Key Takeaways


- Calorimeters measure the change in enthalpy of a given reaction, through detecting a change in temperature of a liquid medium.
- Two main types of calorimeters exist: constant-pressure (coffee-cup) and constant-volume (bomb) calorimeters.
- Specific heat capacity is the amount of energy needed to increase 1 gram of a substance by 1 degree.
- Latent heat capacity is the amount of energy needed to change the state of 1 gram of a substance.

Exercises

Practice Questions



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7.5 Rates of Reaction

Learning Objectives

- Learn about the requirements for a reaction to occur under collision theory: sufficient energy and correct orientation.
- Learn about how reactant concentration, surface area, temperature and the presence of a catalyst can alter the rate of a chemical reaction.

Introduction

Speed plays an important role in many things we do in everyday life. If you sleep in and need to get to your chemistry lecture quickly, you may choose to drive instead of walk because driving is faster. You may use a favourite website to stream video content online because of its quick and reliable download speeds. When applying for a summer job, you may have had to include your typing speed on your resume to show your competency using computers. These examples emphasise that the speed of a process is an important consideration in our everyday lives. Similarly, the speed of a chemical reaction is also a significant consideration and is called its reaction rate. Reaction rates vary dramatically, with some reactions occurring on a time scale of seconds, while other reactions take many thousands of years. Several factors can influence reaction rate, and the study

of the interplay between these factors and the rate of a chemical reaction is called kinetics.

Factors that Affect the Rate of Reactions

Reaction kinetics is the study of the rate of chemical reactions, and reaction rates can vary greatly over a large range of time scales. Some reactions can proceed at explosively fast rates, like the detonation of fireworks (see Figure 7.5.1), while others can occur at a sluggish rate over many years, like the rusting of barbed wire exposed to the elements (see Figure 7.5.2).



Figure 7.5.1: “Fireworks at Night Over River.” The chemical reaction in fireworks happens at an explosive rate. Image attribution: “Fireworks at night over river” © Jon Sullivan is licensed under a Public Domain license.



Figure 7.5.2: “Rusted Barbed Wire.” The rusting of barbed wire occurs over many years. Image attribution: “Barbed wire (rusting after years of hard work)” © 2007 by Waugsberg is licensed under a CC BY-SA (Attribution-ShareAlike) license.

Collision Theory Expanded

Let us expand our understanding of collision theory further. While we have already discussed how temperature increases the kinetic speed of molecules, we must also consider how chemical reactions begin. It is not only enough for molecules to touch — chemical processes only occur when reactant molecules “effectively collide.” For an “effective collision” to occur, the reactant molecules must be oriented in space correctly to facilitate the breaking and forming of bonds and the rearrangement of atoms that result in the formation of product molecules (see Figure 7.5.3).

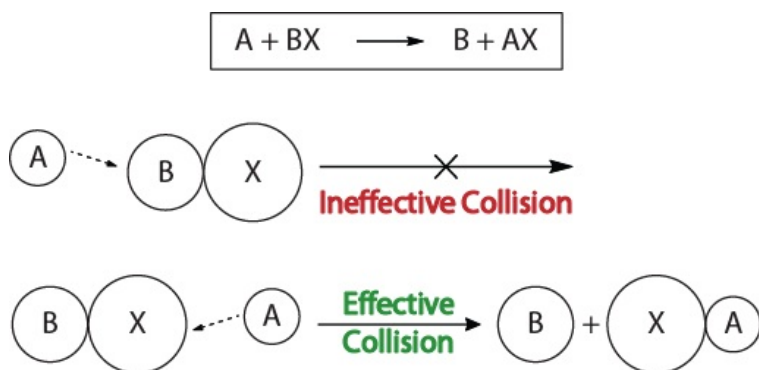


Figure 7.5.3: “Collision Visualisations.” This visualisation shows an ineffective and effective collision based on molecular orientation. Transcript. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

During a molecular collision, molecules must also possess a minimum amount of kinetic energy for an effective collision to occur. This requirement varies for each reaction and is known as the **activation energy (E_a)** (see Figure 7.5.4). The rate of reaction, therefore, depends on the activation energy; a higher activation energy means that fewer molecules will have sufficient energy to undergo an effective collision.

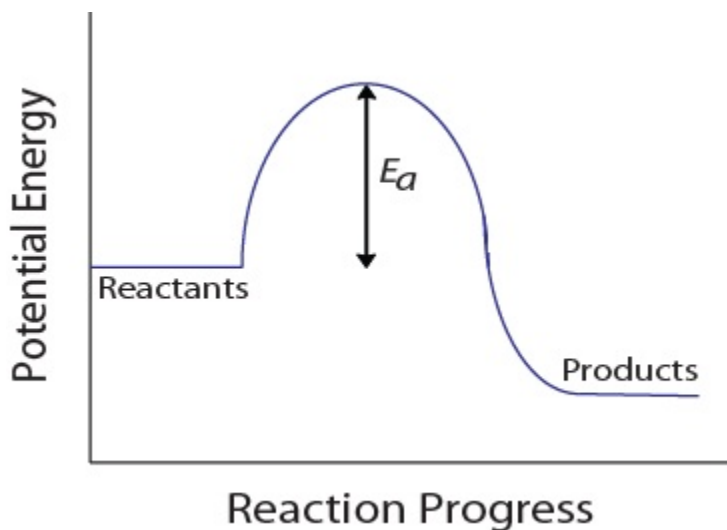


Figure 7.5.4: “Potential Energy and Activation Energy.” This potential energy diagram shows the activation energy of a hypothetical reaction. Transcript. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

Factors That Affect Rate

There are four main factors that can affect the reaction rate of a chemical reaction:

1. **Reactant concentration.** Increasing the concentration of one or more reactants will often increase the rate of reaction. This occurs because a higher concentration of a reactant increases the number of molecules within the same space and will lead to more collisions of that reactant in a specific time period.
2. **Physical state of the reactants and surface area.** If reactant molecules exist in different phases, as in a heterogeneous mixture, the reaction rate will be limited by the surface area of

the phases in contact. For example, if a solid metal reactant and gas reactant are mixed, only the molecules present on the surface of the metal are able to collide with the gas molecules. Therefore, increasing the surface area of the metal by pounding it flat or cutting it into many pieces will increase its reaction rate. This is why sugar cubes dissolve slower than powdered sugar.

3. **Temperature.** An increase in temperature typically increases the rate of reaction. An increase in temperature will raise the average kinetic energy of the reactant molecules. Therefore, a greater proportion of molecules will have the minimum energy necessary for an effective collision (see Figure 7.5.5).

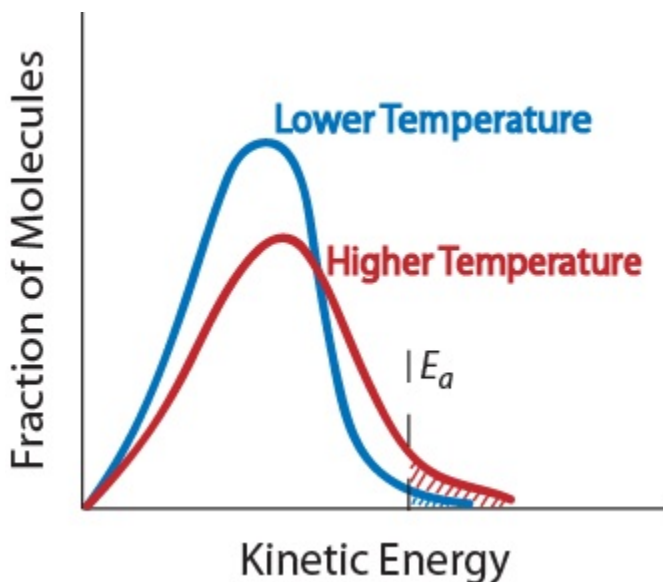


Figure 7.5.5 “Temperature and Reaction Rate.” Effect of temperature on the kinetic energy distribution of molecules in a sample. Transcript. Image attribution: Chem&121: Introduction to Chemistry Copyright © 2023 by Lake Washington Institute of Technology is licensed under a Creative Commons Attribution-NonCommercial-ShareAlike 4.0 International License.

4. **Presence of a catalyst.** A **catalyst** is a substance that accelerates a reaction by participating in it without being consumed. Catalysts provide an alternate reaction pathway to obtain products — that has a lower activation energy than the catalyst-free pathway. They are critical to many biochemical reactions.

Key Takeaways

- Reaction kinetics is the study of the rate of chemical reactions.
- According to collision theory, for a reaction to occur, molecules must collide with one another with sufficient energy and at the correct orientation.
- Collision theory allows us to predict how a variety of factors speed up chemical reactions, such as reactant concentration, surface area, temperature, and catalysts.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=640#h5p-73>

Transcript

PART VIII

CHAPTER 8:
FUNDAMENTALS OF
ORGANIC CHEMISTRY

I. 8.0 What is Organic Chemistry?

What is Organic Chemistry?

Organic chemistry is a branch of chemistry that focuses on the study of carbon-containing compounds. While it initially dealt primarily with compounds derived from living organisms (hence the term “organic”), the definition has evolved to include a vast array of synthetic compounds as well. Carbon is a unique element in that it can form stable and diverse structures through covalent bonding with other carbon atoms and various elements.

Nature is filled with chemical structures of many types, but in the chemistry of life, we find an abundance of organic chemicals. Carbon has an outsized role to play in life and, thus, in our chemical activities. Agriculture, manufactured goods from ag products (like textiles), and pharmaceutical products are all largely based on organic molecules. The petroleum industry and all the products related to that are also linked to organic chemistry because petroleum is a fossil fuel produced in geologic processes from formerly living matter. In modern times, these substances are the raw materials converted into a huge variety of plastics that we use in constructing the built world.

Why Carbon?

	boron 5 B 10.81	carbon 6 C 12.01	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	helium 2 He 4.0026
	aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948
zinc 30 Zn 65.38	gallium 31 Ga 69.723	germanium 32 Ge 72.64	arsenic 33 As 74.922	selenium 34 Se 76.65	bromine 35 Br 79.904	krypton 36 Kr 83.80

Figure 8.0.1: Carbon within the periodic table. Image attribution: Carbon raised within the periodic table © Rob Pitman – stock.adobe.com

Carbon, atomic number 6, is an element with atoms that are small and relatively simple. Its nucleus contains 6 positively charged protons, and there are 6 electrons outside the nucleus distributed into two shells. The outer shell has four electrons that are held quite strongly by the electrostatic pull from the nucleus. So, while a carbon atom can be ionised through either the gain or loss of electrons, it does not tend to do so. Carbon does, however, readily engage in covalent bonding, sharing electrons with neighbouring atoms and forming tight associations with them. The four valence electrons in a carbon atom can do this by forming four single bonds, by forming two single bonds and a double bond, by forming one single bond and a triple bond, or by forming two double bonds.

Carbons also covalently bond with one another, forming chains of

various lengths and rings. It readily bonds with other atoms, such as oxygen, nitrogen and hydrogen, forming quite stable arrangements with these common elements as well.

The architecture of carbon structures is, therefore, fantastically varied. Small organic molecules might contain just one or two carbon atoms surrounded by other atoms. But the larger organic molecules can contain hundreds or thousands of carbons, linked with rings and bridges and other complex structures that fold into particular three-dimensional structures. Figure 8.0.2 shows the chemical structure of the antibiotic amoxicillin, which consists of several carbon atoms forming single and double bonds and rings.

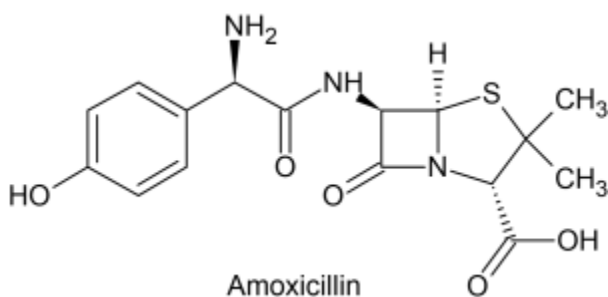


Figure 8.0.2 Amoxicillin.

No other element can quite do what carbon does: silicon has the ability to form four bonds with other atoms, but those bonds tend to be weak due to the additional electron shell in a silicon atom. Nitrogen has five valence electrons, so it generally only forms 3 single bonds, limiting its usefulness. Boron, similarly, does not make for a dependable, stable base structure.

Media Attributions

- Carbon raised within the periodic table © Rob Pitman - stock.adobe.com

8.1 Introduction to Hydrocarbons

Learning Objectives

- Differentiate four classes of hydrocarbons.
- Explain saturated and unsaturated hydrocarbons and how their structures influence their reactivity.

Hydrocarbons are organic compounds that only contain carbon and hydrogen. They are used in a wide variety of applications and on an enormous scale. Hydrocarbons are sources from natural gas and crude oil, the fossil fuel substances taken from underground. These unrefined products are mixtures that are refined in industrial facilities to generate the raw materials used to produce everything from gasoline to plastics, paints, and cosmetics. Hydrocarbons are the simplest organic compounds, but they have interesting physiological effects. These effects depend on the size of the hydrocarbon molecules and the part of the body that is exposed to them.

Hydrocarbons are classified into four main classes: alkanes, alkenes, alkynes, and aromatic hydrocarbons (Figure 8.1.1):

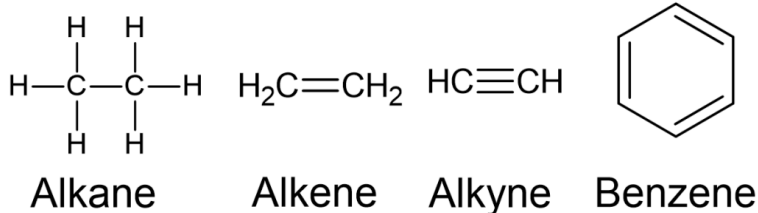


Figure 8.1.1 Different types of hydrocarbons.

Alkanes

- Alkanes are comprised of $\text{C} - \text{C}$ single bonds and have no functional groups.
- They are known as **saturated hydrocarbons** due to the fact that these molecules have the maximum number of hydrogen atoms possible. In alkanes, each carbon atom forms four single covalent bonds, two with adjacent carbon atoms and two with hydrogen atoms. This results in a structure where each carbon is “saturated” with the maximum number of hydrogen atoms it can hold.
- Methane (Figure 8.1.1) is the simplest alkane, and its molecular formula is CH_4 .

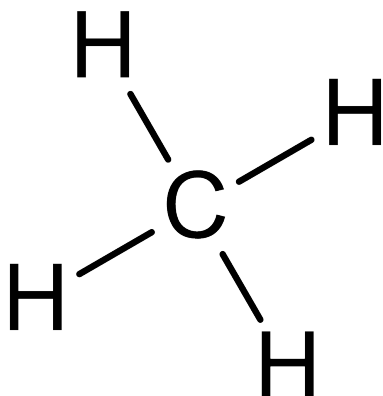


Figure 8.1.2 Structure of methane

- The general formula for alkanes is C_nH_{2n+2} . This formula indicates that for every n carbon atom, there are $2n + 2$ hydrogen atoms. This ensures that each carbon atom is bonded to four other atoms (either other carbon atoms or hydrogen atoms), satisfying the tetravalency of carbon.
- Alkanes can form straight-chain, branched and cyclic structures.
 - **Straight-chain alkanes:** they are distinguished by the arrangement of carbon atoms in a continuous, unbranched chain, as shown in Figure 8.1.3.

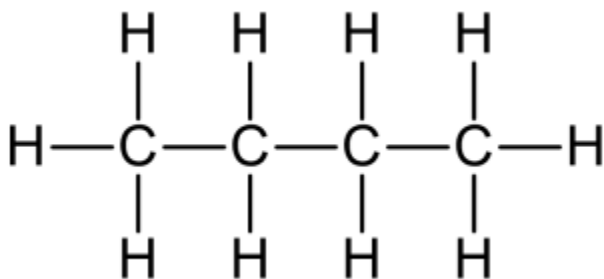


Figure 8.1.3 An example of a straight-chain alkane.

- **Branched alkanes:** alkanes can have branching, meaning that instead of forming a straight chain, some carbon atoms may have additional carbon atoms attached to them, as shown in Figure 8.1.4.

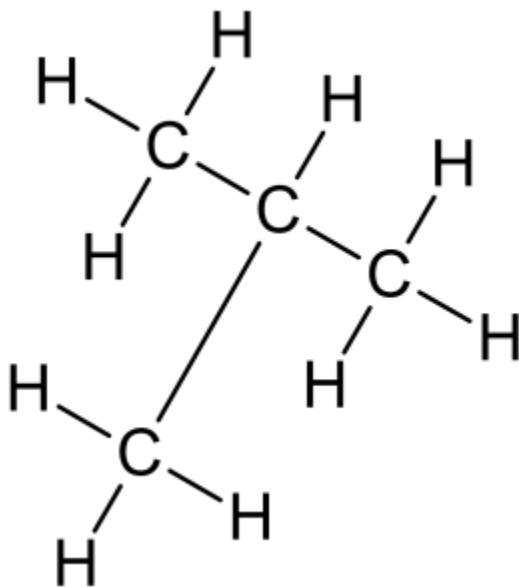


Figure 8.1.4 An example of a branched alkane.

- **Cyclic alkanes/cycloalkanes:** in a cyclic alkane, the carbon atoms are arranged in a ring. The simplest example is cyclopropane C_3H_6 is shown in Figure 8.1.5.

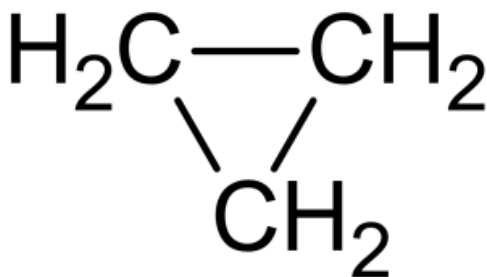


Figure 8.1.5 Structure of cyclopropane.

- Alkanes are known to be less reactive compared to other classes of organic compounds due to:
 - strong single bonds between $\text{C} - \text{C}$ and $\text{C} - \text{H}$ and,
 - the polarity of $\text{C} - \text{C}$ and $\text{C} - \text{H}$ bonds. In alkanes, the electrons in single bonds are approximately evenly distributed between the two bonded atoms, making them non-polar. Therefore, no significant positive or negative charge is present on any part of these molecules to attract other ions or molecules. Hence, alkanes are less susceptible to attack by reagents.
- Despite their general lack of reactivity, alkanes can undergo certain reactions under specific conditions. For example, they can undergo combustion (reaction with oxygen to produce carbon dioxide and water) and can be involved in radical reactions under the influence of high-energy conditions.
- Alkanes of low molar mass—those with from 1 to approximately 10 or so carbon atoms—are gases or light liquids that act as anesthetics. Methane, CH_4 , is an alkane that is the

combustible natural gas you may burn in your furnace to heat your home. Octane, C_8H_{18} , is an alkane that is a component of gasoline. On the skin, liquid alkanes with approximately 5–16 carbon atoms per molecule wash away natural skin oils and cause drying and chapping, while heavier liquid alkanes (those with approximately 17 or more carbon atoms per molecule) act as emollients (skin softeners).

Alkene

- Alkenes are a class of hydrocarbons characterised by the presence of at least one carbon-carbon double bond in their molecular structure.
- These compounds are part of the larger group of unsaturated hydrocarbons, as they contain fewer hydrogen atoms than the corresponding alkanes (saturated hydrocarbons) with the same number of carbon atoms.
- The general molecular formula for alkenes is C_nH_{2n} , where n is the number of carbon atoms. The simplest alkene is ethylene with the molecular formula C_2H_4 (Figure 8.1.6):

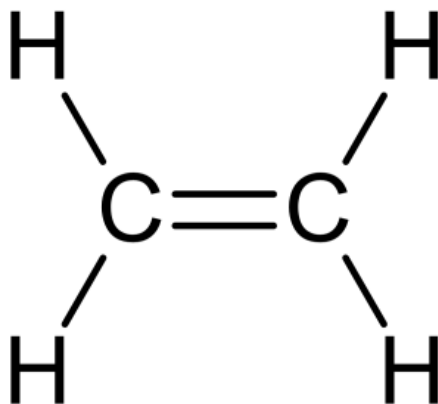


Figure 8.1.6 Structure of ethylene.

- The presence of a double bond introduces a degree of unsaturation into the molecule, allowing alkenes to undergo additional reactions not possible with saturated hydrocarbons.
- The carbon-carbon double bond consists of one sigma (σ) bond, which is formed by the head-on overlap of atomic orbitals, and one pi (π) bond, which is formed by the side-to-side overlap of p orbitals. The presence of the pi bond introduces a degree of rigidity and restricts rotation around the double bond.
- Alkenes are more reactive than alkanes due to the presence of the π bond. The π electrons are more accessible and can participate in various chemical reactions. Alkenes can undergo polymerisation to form polymers. For example, ethene can polymerise to form polyethylene, a widely used plastic.
- Ethene C_2H_4 is a gaseous alkene that serves as a cellular signal in fruits to stimulate ripening. Fruits that are sensitive to this signalling molecule can be placed in a paper bag along with an apple – the apple emits ethene gas, setting off the ripening process in the fruit. Commercial fruit packers can

make use of this phenomenon by harvesting unripe fruits and then inducing ripening right before shipping them to consumers.

- The bonding in alkenes is trigonal planar, and the molecules are unable to rotate along the axis of the bond. The double bonds thus lead to 120-degree bond angles and a planar triangular geometry around the double bond.

Alkynes

- Alkynes are a class of hydrocarbons characterised by the presence of at least one $\text{C} \equiv \text{C}$ triple bond in their molecular structure. Like alkenes, alkynes are unsaturated hydrocarbons, meaning they have fewer hydrogen atoms than the corresponding alkanes with the same number of carbon atoms. The $\text{C} \equiv \text{C}$ triple bond consists of one sigma (σ) bond and two pi (π) bonds. The triple bond imparts a degree of rigidity to the molecular structure.
- The general molecular formula for alkynes is $\text{C}_n\text{H}_{2n-2}$, where n is the number of carbon atoms. The simplest member of the alkyne family is ethyne with the molecular formula C_2H_2 (Figure 8.1.7). The common name of ethyne is acetylene. Acetylene is widely used in oxy-acetylene welding and cutting processes. It is a precursor in the production of polymers. For example, it can polymerise to form polyacetylene, a conductive polymer with potential applications in electronic devices.

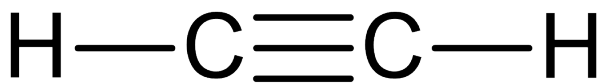


Figure 8.1.7 Structure of acetylene.

- Alkynes are more reactive than alkanes due to the presence of the triple bond and can undergo various chemical reactions like alkenes. The hydrogen atoms attached to the carbon atoms of the triple bond in alkynes are relatively acidic compared to those in alkanes and alkenes. This acidity is a result of the electronegative nature of the adjacent triple bond.
- In alkynes, the geometry around the triple bond is linear (bond angles are 180°), and only one other atom can bond to the alkyne carbon, so there is no rotation.

Aromatic Hydrocarbons

- Aromatic hydrocarbons are a class of organic compounds that contain a cyclic structure with alternating single and double bonds, known as an aromatic ring. In the aromatic ring, each carbon atom is bonded to two other carbon atoms and one hydrogen atom. The electrons in the pi (π) bonds are delocalised, creating a stable, resonant structure.
- Aromatic hydrocarbons are unsaturated, like alkenes and alkynes. Although aromatic hydrocarbons contain double bonds, their reactivity differs from that of typical alkenes. Aromatic compounds are known for their stability and resistance to certain reactions that would break the aromaticity of the ring.
- Benzene (C_6H_6) is the simplest and most well-known aromatic hydrocarbon (Figure 8.1.8). Other common aromatic hydrocarbons include toluene, xylene, naphthalene, and anthracene. Aromatic rings can also be found in many organic compounds, including certain classes of drugs, dyes, and natural products.

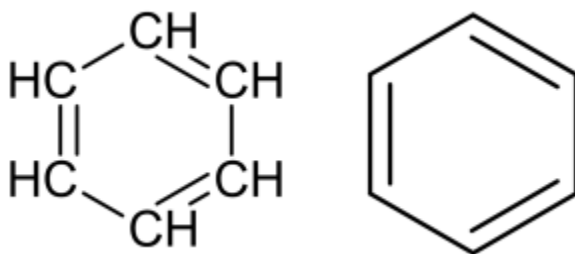


Figure 8.1.8 Benzene

- Aromatic hydrocarbons have numerous industrial applications. They are key components in the production of plastics, synthetic fibres, detergents, pharmaceuticals, and many other chemicals.
- Aromatic groups are planar (flat) ring structures and are widespread in nature, so you will see them frequently if you encounter chemical structures in biology classes or in biomedical work.

Key Takeaways

- Hydrocarbons are organic compounds that only contain carbon and hydrogen.
- Hydrocarbons are classified into four main classes: alkanes, alkenes, alkynes, and aromatic hydrocarbons.
- Alkanes are comprised of $\text{C} - \text{C}$ single bonds, alkenes consist of one or more $\text{C} = \text{C}$ double bond and alkynes are made up of one or more $\text{C} \equiv \text{C}$

triple bonds.

- Alkanes are classified as saturated hydrocarbons, while alkenes, alkynes and aromatic hydrocarbons are classified as unsaturated hydrocarbons.
- Alkanes are less reactive compared to other hydrocarbons due to the presence of strong $\text{C} - \text{C}$ and $\text{C} - \text{H}$ single bonds.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=2109#h5p-67>

Transcript

8.2 Drawing and Interpreting Organic Formulas

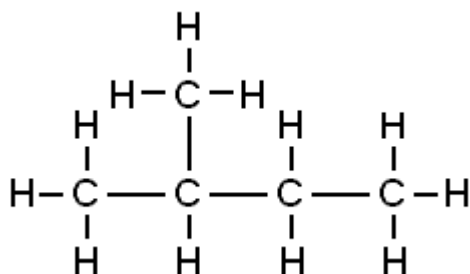
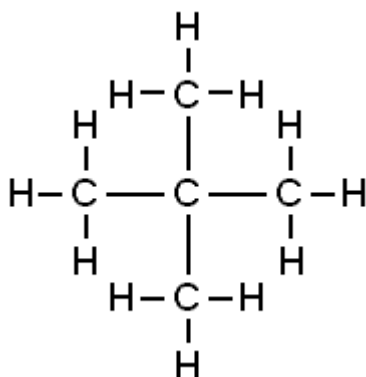
Learning Objectives

- Understand different ways of presenting organic molecules.
- Comprehend different bonding patterns in organic molecules.
- Draw and interpret organic molecules using structural formulas, condensed structural formulas and line bond structures.

Drawing Conventions

Organic molecules can be large and can contain multiple atoms of carbon, hydrogen and other elements. Their architecture can be complicated, with chains of various lengths and ring structures.

Consider the structures shown here:



Both of these structures are representations of organic molecules. Each contains 5 carbon atoms and 12 hydrogen atoms. But they look different, and the properties of these substances are similar but not exactly the same.

The substance to the left has a boiling point of 9.5 °C, making it a gas at room temperature. The substance to the right has a boiling point of 28 °C, so it is a liquid at room temperature.

As we begin to consider and study organic molecules, we are very quickly faced with this fact: molecular structures (showing the number and types of atoms in a molecule, such as H_2O) are inadequate to describe these substances.

In organic chemistry, we make use of different types of formulas for this reason. There are several types, including:

- **Structural formulas:** which show every atom with its elemental symbol and every bond drawn as a line. These look a

lot like the Lewis Structures you learned to draw in Chapter 3, but without the nonbonding electron pairs. Figure 8.2.1 shows an example of a structural formula.

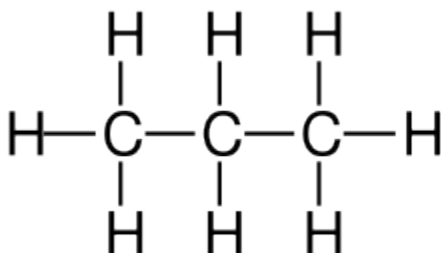


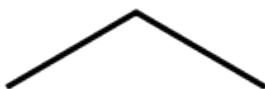
Figure 8.2.1 An example of a molecule drawn using the structural formula.

- **Condensed structural formulas:** similar to structural formulas but without bonds shown to hydrogen atoms, so a carbon with 3 hydrogens attached becomes a CH_3 . There are various levels of condensed structural formulas, and there are some rules about how formulas are properly condensed. The condensed formula of the molecule comprised of three carbons and eight hydrogens is shown below.



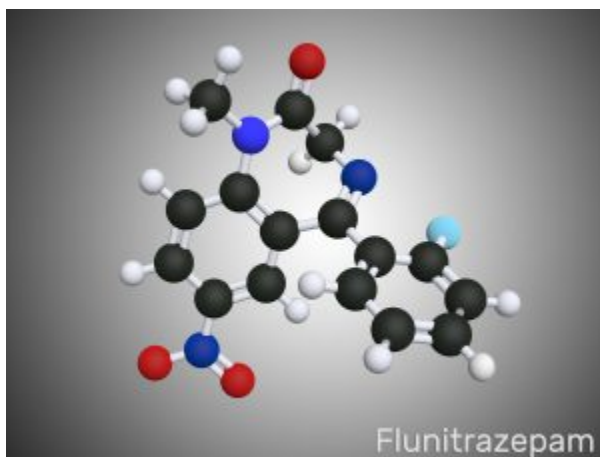
- **Line-bond, also called skeletal structures:** these are the hardest to learn but the quickest to draw among those described here. They are used extensively in chemical communication. In line-bond structures, all bonds between carbons and between carbons and other atoms except hydrogen are shown as displayed in Figure 8.2.2. Hydrogens are not included unless they are attached to something other than

carbon. The carbon atoms are shown without the elemental symbol. Instead, the viewer recognises carbon as present anywhere there is a vertex (pointed place) in the drawing and at the end of a line.

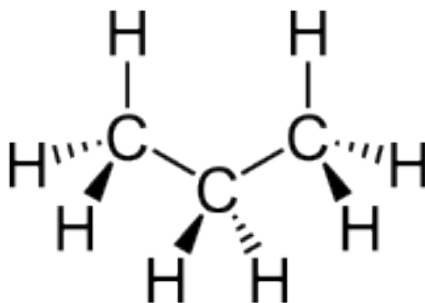


8.2.2 An example of a molecule drawn using the line-bond structure.

- **Ball-and-stick model:** the ball-and-stick model is a molecular model used in chemistry to represent the three-dimensional arrangement of atoms in a molecule. In this model, atoms are represented by spheres (balls), and chemical bonds between atoms are represented by sticks or rods, as shown in Figure 8.2.3. This helps to visualise the arrangement of atoms in a molecule and gain insights into the molecular geometry, bond angles, and overall structure.



- **Dash-wedge structures:** these structures are a common way to represent the three-dimensional arrangement of atoms in a molecule on a two-dimensional surface, such as paper or a computer screen. These structures use lines (dashes and wedges) to indicate the orientation of bonds in three-dimensional space, as shown in Figure 8.2.4. A dash is used to represent a bond that extends away from the viewer, going into the plane of the paper or screen. It suggests that the atom at the end of the bond is situated behind the plane of the paper. A wedge is used to represent a bond that comes out of the plane of the paper or screen toward the viewer. It suggests that the atom at the end of the bond is closer to the viewer.



8.2.4 An example of a molecule drawn using the dash-wedge structure.

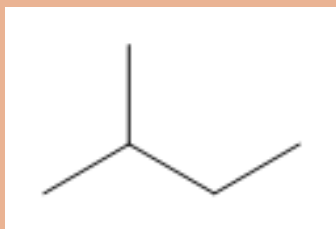
Examples 8.2.1

Test Yourself

Can you tell which of the two structures is shown in the examples here?



b)



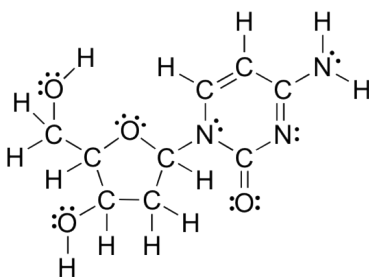
Answers

- a) Condensed structural formula
- b) Line bond/skeletal structures

Common bonding patterns in organic structures

Drawing structural formulas is a good starting point for a novice organic chemist. This works when dealing with small, simple structures, but when you start dealing with larger structures, it becomes increasingly difficult and time-consuming. Imagine trying to draw all atoms and bond every time you want to discuss the structure below (Figure 8.2.5), which is one small piece of DNA:

Large molecules such as this are commonly considered in organic chemistry and biochemistry. In these situations, line-bond structures really help. To get good at drawing them accurately, you will first want to get familiar with some common bonding arrangements involving elements frequently found in organic molecules.



2'-deoxycytidine

Figure 8.2.5 Structural formula of 2'-deoxycytidine

- Let's start with carbon. Carbon is said to be **tetravalent**, meaning that it tends to form four bonds. If you look at a variety of structures, including carbon, you can see that nearly always, each carbon atom has four bonding pairs of electrons, each represented as a line surrounding it (Figure 8.2.6).

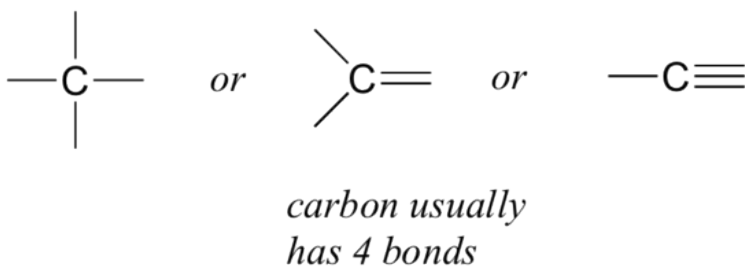


Figure 8.2.6 Common bonding patterns of carbon.

This is a pattern that holds throughout most of the organic molecules we will see.

If carbon has other electron arrangements in its valence shell (in other words, if it does not fulfil the octet rule), it will

have a formal charge or exist as a radical, as shown in Figure 8.2.8:

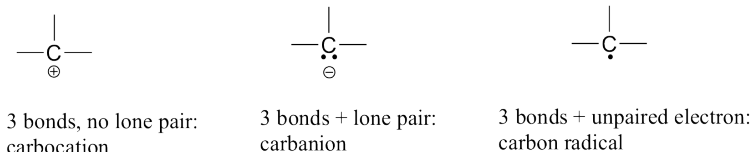


Figure 8.2.8 Bonding patterns of carbon when it does not obey the octet rule.

If you are able to quickly recognise these patterns (and the patterns described below for other atoms), it will help you tremendously as you learn more about organic chemistry.

- The pattern for hydrogen is easy: hydrogen atoms have only one bond and no formal charge. As a rule, all hydrogen atoms in organic molecules have one bond and no formal charge.
- For oxygen, you will see the atom bonding in three ways, as shown in Figure 8.2.9, all of which fulfil the octet rule.

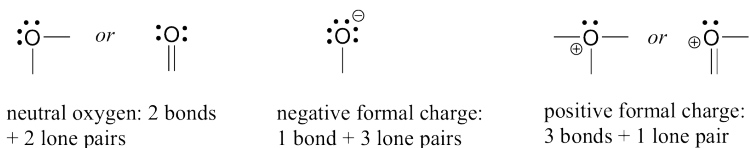


Figure 8.2.9 Common bonding patterns of oxygen.

In most cases, an oxygen atom has two bonds and two lone pairs, as it does in water. In this arrangement, it will have a formal charge of zero. If it has one bond and three lone pairs, as in the hydroxide ion, it will have a formal charge of -1. If

it has three bonds and one lone pair, it will have a formal charge of +1.

There are, again, some additional possibilities. However, these three examples will account for virtually everything we see.

- Nitrogen has two major bonding patterns, both of which fulfil the octet rule, as shown below in Figure 8.2.10:

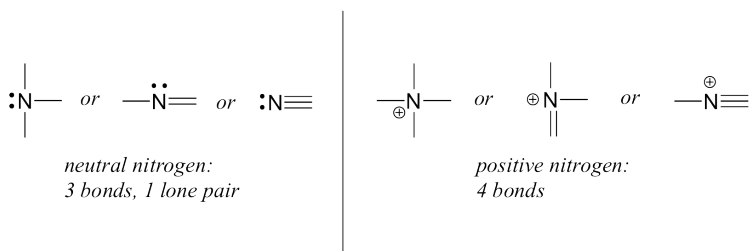
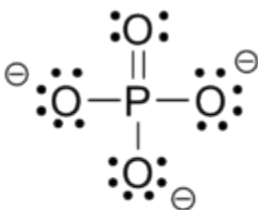


Figure 8.2.10 Common bonding patterns of nitrogen.

If nitrogen has three bonds and a lone pair, it has a formal charge of zero. If it has four bonds (and no lone pair), it has a formal charge of +1. In a fairly uncommon bonding pattern, negatively charged nitrogen has two bonds and two lone pairs.

- The third-row elements are commonly found in important organic molecules: sulphur and phosphorus. Although both of these elements have other bonding patterns that are relevant in laboratory chemistry, in a biological context, sulphur almost always follows the same bonding/formal charge pattern as oxygen, while phosphorus is present in the form of phosphate ion (PO_4^{3-}) (see Figure 8.2.11), where it has five bonds (almost always to oxygen), no lone pairs, and a formal charge of zero. Remember that atoms of elements in the third row and

below in the periodic table have 'expanded valence shells' with *d* orbitals available for bonding, and the octet rule does not always apply.



phosphate

Figure 8.2.11 One of the bonding patterns of phosphate.

- Finally, the halogens (fluorine, chlorine, bromine, and iodine) are very important in laboratory and medicinal organic chemistry but are less common in naturally occurring organic molecules. Halogens in organic compounds are usually seen with one bond, three lone pairs, and a formal charge of zero. Sometimes, especially in the case of bromine, we will encounter reactive species in which the halogen has two bonds (usually in a three-membered ring), two lone pairs, and a formal charge of +1. Common bonding patterns of halogens are shown below in Figure 8.2.12.

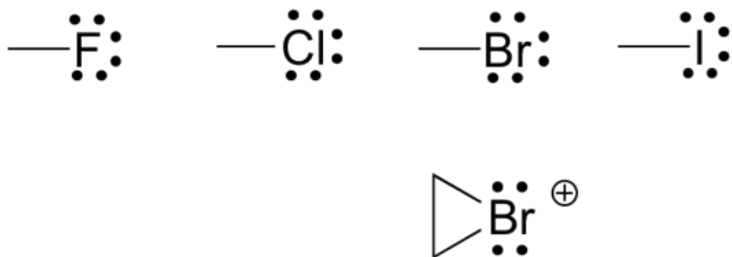


Figure 8.2.12 Common bonding patterns of halogens.

Organic structure types often do not include lone pairs since you can assume that the proper number of electrons are present around each atom to match the indicated formal charge (or lack thereof). Occasionally, though, lone pairs are drawn if doing so helps to make an explanation more clear.

Using the ‘Line Bond Structure’ Convention

In organic chemistry, the way compounds are drawn is somewhat different from the Lewis Structures you are used to seeing in your general chemistry. In some sources, you will see **condensed structures** for smaller molecules instead of full structural formulas that include every bond, as shown in Figure 8.2.13.

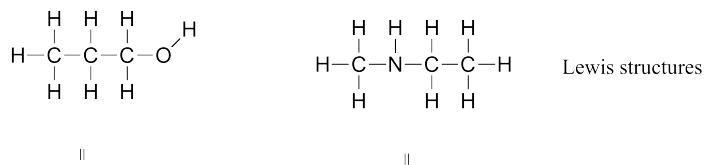


Figure 8.2.13 Lewis structures and condensed structures in organic chemistry.

More commonly, organic and biological chemists use an abbreviated drawing convention called **line-bond structures**, also sometimes called skeletal structures. The convention makes it easier to draw molecules, but the convention does need to be learned.

Some considerations are:

- Carbon atoms are not depicted with their elemental symbol but rather by a vertex (corner) or a free end of a bond.
- Open-chain molecules are usually drawn out in a 'zig-zig' shape.
- Hydrogens attached to carbons are generally not shown; rather, like lone pairs, it is assumed the person viewing the structure knows where they are.
- Hydrogens bonded to nitrogen, oxygen, sulphur, or anything other than carbon *are* shown but are usually drawn without showing the bond.

The following examples illustrate the convention (Figure 8.2.14).

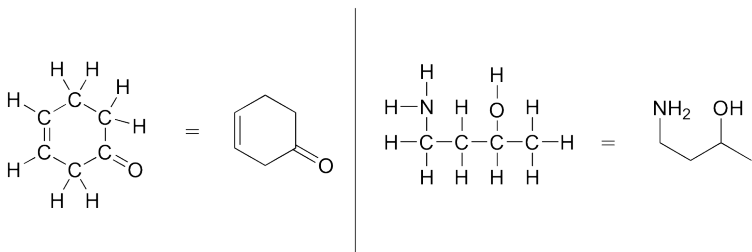


Figure 8.2.14 Examples of molecules drawn using line-bond structures.

Line-bond structures make it much easier to see the basic structure of the molecule and the locations where there is something other than **C—C** and **C—H** single bonds. It can also be drawn quickly.

Sometimes, one or more carbon atoms in a line structure will be depicted with a capital C if doing so makes an explanation easier

to follow. If you label a carbon with a C, you also must draw in the hydrogens for that carbon.

Key Takeaways

- Organic molecules can be represented in different ways.
- Understanding the common bonding patterns of common atoms found in organic chemistry is essential for drawing and interpreting organic structures.
- The line bond structure is the commonly used method to represent organic molecules.

Exercises

Practice Questions

A good way to test your understanding of the line structure convention is to see if you can accurately determine the number of hydrogen atoms in a molecule from its line structure. Do this for the structures here.



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How many hydrogens are bonded to carbon at each hotspot (identified with a purple plus sign) on the structure shown below? Click on the purple spot to find the correct answer.



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Transcript

Media Attributions

- Flunitrazepam drug molecule. It has hypnotic, sedative, anxiolytic properties. Molecular model © bacsica-stock.adobe.com

8.3 Naming Organic Compounds

Learning Objectives

- Explain the importance of having a standard naming system in organic chemistry.
- Understand the main guidelines for naming an organic compound using IUPAC nomenclature.

Chemists commonly use drawings and names to communicate about organic structures. While names are preferred in certain situations, the vast number of known organic substances makes naming each one independently impractical. To address this, the International Union of Pure and Applied Chemistry (IUPAC, usually pronounced **eye-you-pack**) has developed a naming system. This system allows for the description of a chemical by components of a name, providing enough information for a complete and unique structure. Although the IUPAC system is frequently used for simpler molecules, more complex substances, especially those produced by biological organisms, may have lengthy names.

How to get a systematic name from a structure

- To assign a name to a compound, begin by determining the

parent chain, which is the longest straight chain of carbon atoms. We'll start by considering the simplest straight-chain alkane structures.

If the parent chain is just one carbon long, the name is based on CH_4 , which is called **methane**. For a two-carbon parent chain, the name will be based on C_2H_6 , which is **ethane**. Figure 8.3.1 displays the prefixes used in naming hydrocarbons (1, 2, 3, etc., means the number of carbons present in the longest continuous carbon chain.)

Hydrocarbon Prefixes

1	Meth-
2	Eth-
3	Prop-
4	But-
5	Pent-
6	Hex-
7	Hept-
8	Oct-
9	Non-
10	Dec-

Figure 8.3.1 Hydrocarbon prefixes. Image attribution: AdobeStock_519618928 © Zizo – stock.adobe.com.

- Substituents branching from the main parent chain are given a location signifier. This is done by providing the counted carbon number within the parent chain where the branch exists, with the *lowest possible numbers* being used. For example, notice

below (Figure 8.3.2) how the compound on the left is named 1-chlorobutane, not 4-chlorobutane. “1” designates the chlorine is attached to the first carbon in the parent chain. When the substituents are small carbon-containing, so-called alkyl groups, the terms **methyl**, **ethyl**, and **propyl** are used to identify them.

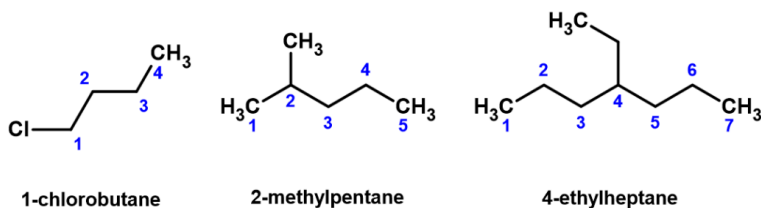
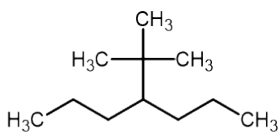
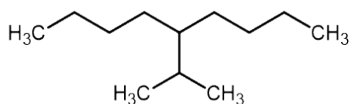


Figure 8.3.2 Line-bond structures of 1-chlorobutane, 2-methylpentane, and 4-ethylheptane.

Other common names for more complex alkyl groups are isopropyl, tert-butyl, and phenyl, as shown in Figure 8.3.3. You may recognise how complicated the names could become, with multiple branches and non-carbon substituent groups all possible on large chains, etc. In some situations, this has caused a preference for common names to be used in casual talk or even among scientists, such as the names given for the amino acids shown in Figure 8.3.3. Some common names, such as phenylalanine, include components of systematic names within them. No one can learn all the common names, and no one can learn all the rules for systematic names in a short period of time. For now, we are learning bits and pieces, and learning how the system for nomenclature works.

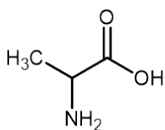


4-*tert*-butylheptane

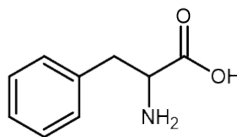


5-isopropylnonane

non-IUPAC examples



Alanine (an amino acid)

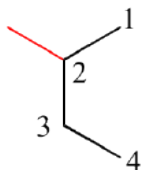


Phenylalanine (an amino acid)

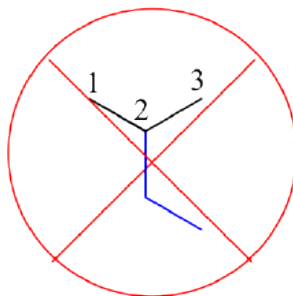
Figure 8.3.3 Line-bond structures of amino acids ‘alanine’ and ‘phenylalanine’ are shown with their IUPAC names.

- The structure shown in Figure 8.3.4 is laid out on the page so that the longest continual carbon chain is oriented vertically. Structures that are presented this way can be confusing, leading to misinterpretation. In this case, the structure could be accidentally named 2-ethylpropane (incorrect) instead of 2-methylbutane (correct).

Keep in mind the IUPAC name for straight-chain hydrocarbons is always based on the **longest possible parent chain**, which in this case is four carbons, not three. Especially if you are looking at large and complicated structures, it can get tricky to identify the parent chain, but it is the foundation of the name.



2-methylbutane



not 2-ethylpropane

Figure 8.3.4 Line-bond structure of 2-methylbutane.

When carbons bond to form rings, the resulting cyclic alkanes are called cyclopropane, cyclobutane, cyclopentane, cyclohexane, and so on (Figure 8.3.5):

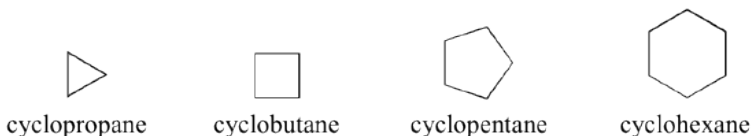
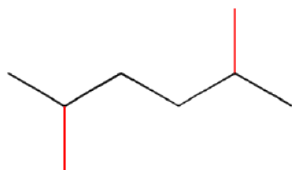


Figure 8.3.5 Line-bond structures of cyclopropane, cyclobutane, cyclopentane and cyclohexane

- In cases where multiple copies of the same substituent are on a structure, the prefixes *di*, *tri*, and *tetra* are used. For instance, if there are two methyl groups present, we name it 'dimethyl', and the presence of three fluoride atoms will be named 'trifluoro' (Figure 8.3.6).



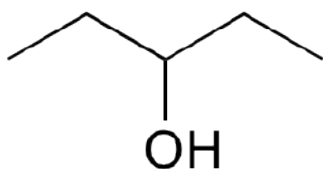
2,5-dimethylhexane



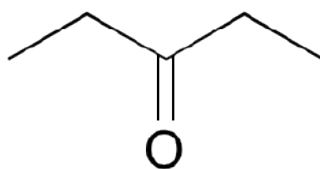
1,2,2-trifluoropentane

Figure 8.3.6 Line-bond structures of 2,5-dimethylhexane and 1,2,2-trifluoropentane.

- We will learn more about functional groups soon. But for now, it is important to highlight that these recognisable groups of atoms show up in names as characteristic suffixes. Alcohols, for example, have ‘**ol**’ appended to the parent chain name, along with a number designating the location of the alcohol group (Figure 8.3.7). Ketones are a functional group with a double bond to oxygen, designated in names by the suffix ‘**one**’ (Figure 8.3.8). For alkanes, the suffix would be ‘**ane**’.



3-pentanol

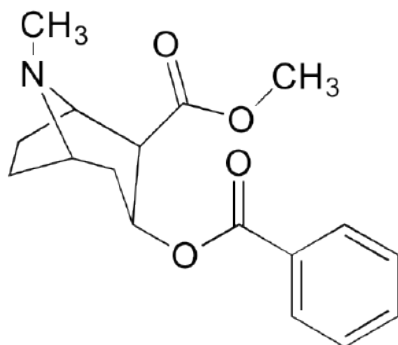


3-pentanone

Figure 8.3.7 Line-bond structures of 3-pentanol and 3-pentanone.

- All of the examples we have seen so far have been simple in the sense that only one functional group was present on each molecule. There are, of course, many more rules in the IUPAC

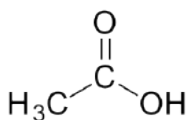
system, and as you can imagine, the IUPAC naming of larger molecules with multiple functional groups, ring structures, and substituents can get very unwieldy very quickly. The drug cocaine shown below in Figure 8.3.8, for example, has the IUPAC name ‘methyl (1*R*,2*R*,3*S*,5*S*)-3-(benzoyloxy)-8-methyl-8-azabicyclo[3.2.1] octane-2-carboxylate.’



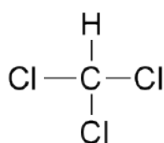
cocaine
(methyl (1*R*,2*R*,3*S*,5*S*)-3-(benzoyloxy)-8-methyl-8-azabicyclo[3.2.1] octane-2-carboxylate)

Figure 8.3.8 Line-bond structure of cocaine.

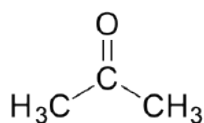
You can see why the IUPAC system is not used very much in biological organic chemistry – the molecules are just too big and complex. A further complication is that, even outside of a biological context, many simple organic molecules are known almost universally by their ‘common’, rather than IUPAC names. The compounds acetic acid, chloroform, and acetone are only a few examples (Figure 8.3.9).



acetic acid



chloroform



acetone

Figure 8.3.9 Structures of acetic acid, chloroform and acetone.

Key Takeaways

- The IUPAC system is used in naming organic compounds.
- The prefix is based on the number of carbon atoms present in the parent chain.
- The suffix is selected based on the functional group present in the structure.
- Substituents present on the structure should be located and named.
- Biological compounds can have complicated names due to the complexity of their structures.

Exercises

Practice Questions



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Transcript



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8.4 Naming and Drawing Hydrocarbons

Learning Objectives

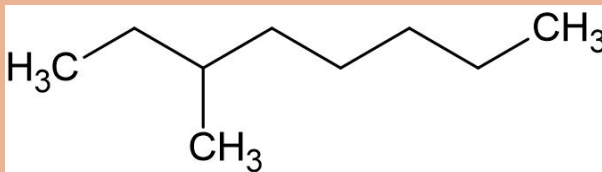
- Name hydrocarbons using IUPAC nomenclature.
- Draw simple hydrocarbons.

Naming Hydrocarbons

Let's apply what we have learned in section 8.3 to name an alkane.

Examples 8.4.1

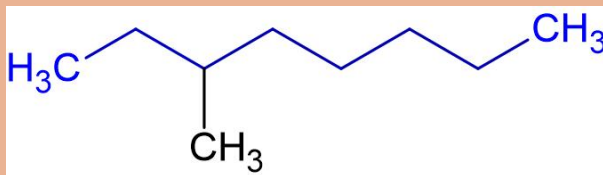
Write the IUPAC name for the following molecule:



Step 1: Identify and name the parent carbon chain.

Find the longest continuous chain of carbon atoms in the molecule. This chain is the main carbon backbone of the alkane. The name of the alkane is based on the number of carbon atoms in the longest chain. Use appropriate prefixes to name the carbon chain. Then, add the suffix 'ane' to indicate membership in the alkane family.

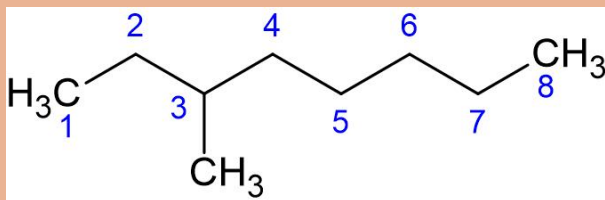
For the example shown below, the longest carbon chain/parent chain is written in one line (highlighted in yellow). However, this might not be the case for some molecules. Then you might need to turn corners to search for it. As the given molecule contains 8 carbon atoms in its parent chain, it can be named an octane ('Oct' is the prefix, and 'ane' is the suffix).



Step 2: Number the carbon atoms in the parent chain.

Number the carbon atoms in the main chain starting from the end nearest to the first substituent (if any). Assign the lowest number possible to the substituents.

The following molecule has one substituent. We should number the carbon atoms, giving the lowest possible number to the carbon atom attached to the substituent. When we start numbering from the left side, substituted carbon gets the third position. If we start numbering from the right side, substituted carbon gets the sixth position. Therefore, we start numbering from the left as shown in the following figure.

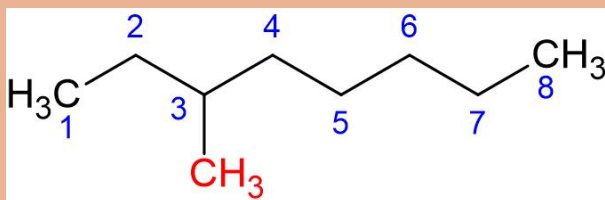


Step 3: Name and locate substituents.

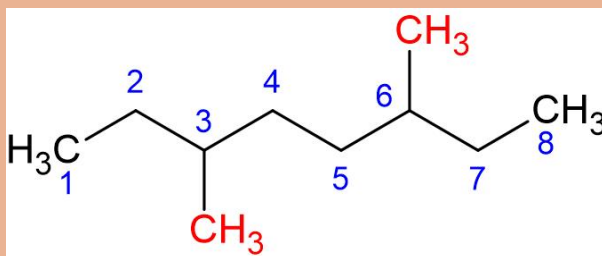
If there are any substituents (groups attached to the main chain), name them using prefixes like methyl, ethyl, propyl, etc. or functional group name, for instance, hydroxy (for OH⁻, chloro (Cl⁻ etc.

Specify the location of each substituent by indicating the number of the carbon atom to which it is attached.

The following molecule has a methyl substituent CH₃ attached to the third carbon in the parent chain.



Sometimes, there can be more substituents attached to different carbons or the same carbon. For instance, consider the following structure:



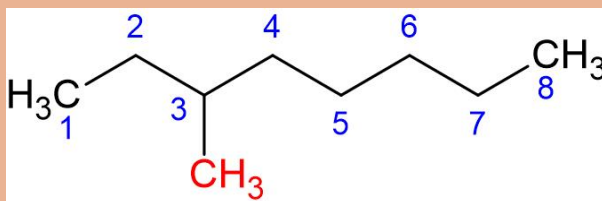
This structure has two methyl groups attached to the third and sixth carbons.

Step 4: Complete the name.

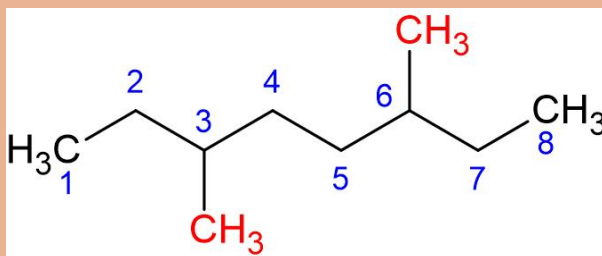
Write the complete name in one word, using hyphens to separate words and commas to separate numbers.

Alphabetise the substituent names when there are multiple substituents. Use appropriate prefixes, such as di, tri, etc., if the same substituent is present more than once.

The IUPAC name of the following molecule would be 3-Methyloctane:



The IUPAC name of the following molecule would be 3,6-Dimethyloctane:



Naming Alkenes and Alkynes

When faced with a structure containing a functional group, such as an alkene, the name of the related alkene can be a good starting point. Most elements of the name will be the same, with the exception that the identity and location of the functional group itself need to be conveyed somehow. For the alkenes, the suffix used is no longer 'ane' but is now 'ene.' The location of the double bond is identified with a number. Count a parent chain that includes the alkene, counting from the end of the chain with the lowest possible number assignment given to the double bond. Then, use the number of the carbon where the double bond is first encountered as the location indicator. Current IUPAC rules put the number immediately before the 'ene' suffix, but name changes are sometimes accepted rather slowly; it remains very common to see this number earlier in the name.

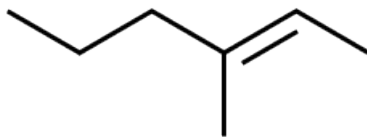


Figure 8.4.1 Structure of
3-methyl-hex-2-ene

Acceptable names for the above molecule (Figure 8.4.1) include 3-methylhex-2-ene and 3-methyl-2-hexene. IUPAC rules encourage placing the location identifier close to the feature at that location. The first name follows IUPAC rules to the letter. However these names can seem awkward even to chemists, and the second form is used frequently.

Other aspects of naming alkenes are identical to the process used for alkanes: the parent chain is indicated by the base name, and the branches are numbered and named just as they are for alkanes.

Alkynes are named similarly to alkenes.

Summary of Naming Rules for Alkenes and Alkynes

1. The longest chain of carbon atoms *containing the double or triple bond* is considered the parent chain. It is named using the same stem as the alkane having the same number of

carbon atoms but ends in *-ene* to identify it as an alkene. Thus, the compound $\text{CH}_2=\text{CHCH}_3$ is *propene*. Alkynes are similarly indicated, using the suffix *-yne*.

2. If there are four or more carbon atoms in a chain, we must indicate the position of the double or triple bond. The carbon atoms are numbered so that the first of the two that are doubly or triply bonded is given the lower of the two possible numbers. The compound $\text{CH}_3\text{CH}=\text{CHCH}_2\text{CH}_3$, for example, has a double bond between the second and third carbon atoms. Its name is 2-pentene (not 3-pentene).
3. Substituent groups are named as with alkanes, and their position is indicated by a number. Thus, Figure 8.4.2 is 5-methyl-2-hexene. Note that the numbering of the parent chain is always done in such a way as to give the double bond the lowest number, even if that causes a substituent to have a higher number. *The double bond always has priority in numbering.*

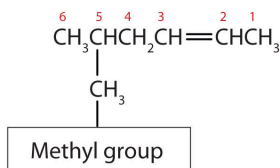
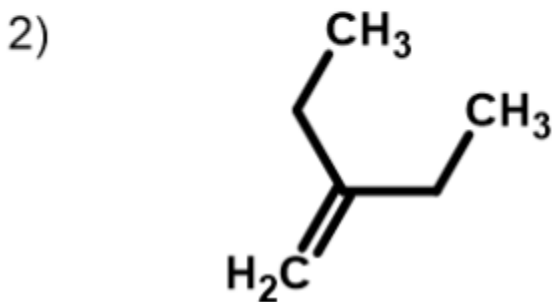
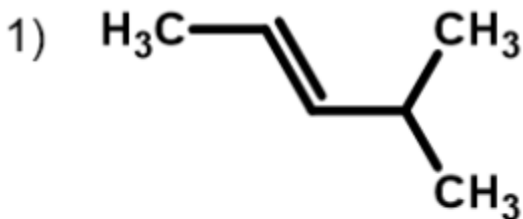


Figure 8.4.2 Structure of 5-methyl-2-hexene

Examples 8.4.2

Problem

Name each of the following compounds.



Solution

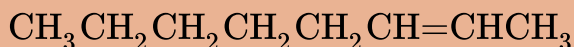
1. The longest chain containing the double bond has five carbon atoms, so the compound is a *pentene* (rule 1). To give the first carbon atom of the double bond the lowest number (rule 2), we number from the left,

so the compound is a 2-pentene. There is a methyl group on the fourth carbon atom (rule 3), so the compound's name is 4-methyl-2-pentene.

2. The longest chain containing the double bond has four carbon atoms, so the parent compound is a *butene* (rule 1). (The longest chain overall has five carbon atoms, but it does not contain the double bond, so the parent name is not *pentene*.) To give the first carbon atom of the double bond the lowest number (rule 2), we number from the left, so the compound is a 1-butene. There is an ethyl group on the second carbon atom (rule 3), so the compound's name is 2-ethyl-1-butene.

Test Yourself

Name the following compound:



Answer

2-Octene

Drawing Hydrocarbons

Let's apply what we have learned about drawing organic molecules in section 8.2 to draw hydrocarbons.

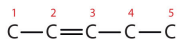
Examples 8.4.3

Draw the structure for each compound.

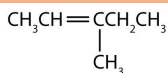
1. 3-methyl-2-pentene
2. cyclohexene

Solution

1. First, write the parent chain of five carbon atoms: C—C—C—C—C. Then add the double bond between the second and third carbon atoms:



Now, place the methyl group on the third carbon atom and add enough hydrogen atoms to give each carbon atom a total of four bonds:



2. Just as there are cycloalkanes, there are *cycloalkenes*. These compounds are named like alkenes but with the prefix *cyclo-* attached to the beginning of the parent alkene name. First, consider what each of the three parts of the name means. *Cyclo* means a ring compound, *hex* means 6 carbon atoms, and *-ene* means a double bond:



Key Takeaways

- All hydrocarbons are named in a similar way; however, the position of double and triple bonds present in alkenes and alkynes, respectively, need to be stated in the name.
- IUPAC name of an organic molecule provides sufficient information to draw the respective molecule.
- Hydrocarbons can be drawn using line-bond and structural formulas.

Exercises

Practice Questions



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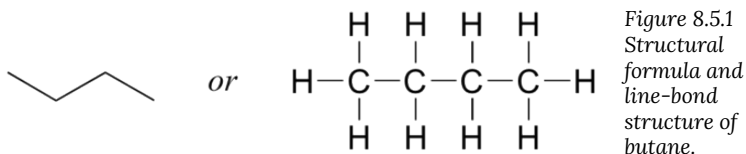
8.5 Isomers

Learning Objectives

- Identify and draw the constitutional isomers.

Constitutional Isomers

Imagine you were asked to draw a structure for a compound with the molecular formula C_4H_{10} . This would not be difficult – you could simply draw as shown Figure 8.5.1:



But when you compared your answer with that of a classmate, they may have drawn this structure as shown below:

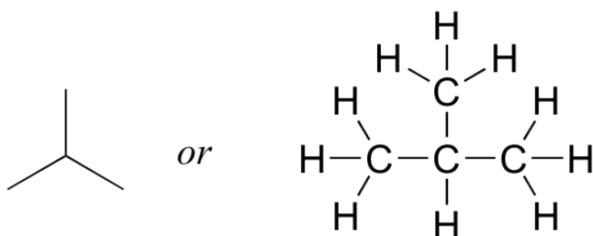


Figure 8.5.2
Structural
formula and
line-bond
structure of
butane.

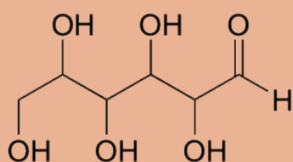
Who is correct? The answer, of course, is that both of you are. A molecular formula only tells you *how many* atoms of each element are present in the compound, not what the actual atom-to-atom connectivity is. There are often many different possible structures for one molecular formula.

Compounds with this relationship, the same molecular formula but different connectivity, are called **constitutional isomers** (sometimes the term '**structural isomer**' is also used). The Greek term 'iso' means 'same.'

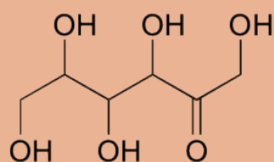
The word isomer is a relational word that describes the relationship between two things. It is similar in that sense to the word 'cousin' or 'uncle.' Being a cousin is not an identity essential to you but describes your relationship to someone else. A substance can be an isomer to *something else*, but being an isomer is not an essential internal characteristic.

Examples

Fructose and glucose, two kinds of sugar molecules, are related as constitutional isomers of one another. Can you figure out the molecular formula for each?

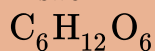


glucose



fructose

Answer



Key Takeaways

- Constitutional isomers, also known as structural isomers, are compounds that have the same molecular formula but differ in the way the atoms are connected or arranged.

Exercises

Practice Questions



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Transcript

8.6 Heteroatoms and Functional Groups

Learning Objectives

- Identify the different functional groups present in organic molecules.

Functional groups in organic compounds

Functional groups are structural units within organic compounds defined by specific bonding arrangements between specific atoms. Many, but not all, functional groups contain **heteroatoms**: atoms other than carbon and hydrogen. The structure below (Figure 8.6.1) of capsaicin, the heat-sensation-producing molecule in hot peppers, incorporates several functional groups, labelled in Figure 8.6.1 and explained throughout this section.

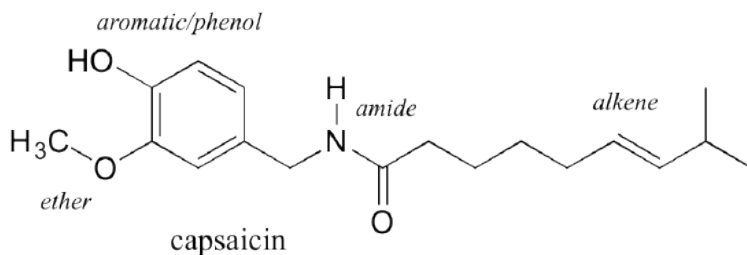


Figure 8.6.1 Structure of Capsaicin.

Functional groups are the key structural elements that define how organic molecules act. Our focus for now will be on drawing and recognising each functional group, as depicted by structural formulas and line-bond structures. But as the name implies, functional groups are linked to the behaviour of substances and their impact on the physical and chemical properties of substances. As the structural feature of a wing on an animal is associated with its ability to fly, functional groups on molecules are structural features associated with what those substances can do.

This section includes a quick tour through a collection of functional groups. You are not expected to know all the details entirely after the first read-through. The overview approach can help you appreciate the variety of structures within organic chemistry and can help you begin to build a vocabulary:

- **Alkanes:** The 'default' in organic chemistry (essentially, the *lack* of any functional groups) is described as an alkane, characterised by single bonds between carbon and carbon or between carbon and hydrogen.
- **Alkenes and alkynes:** Alkenes (sometimes called olefins) have carbon-carbon double bonds, and **alkynes** have carbon-carbon triple bonds.
- **Aromatic hydrocarbons:** The **aromatic** group is exemplified by benzene and naphthalene.

Further details about hydrocarbons can be found in section 8.1.

- **Alkyl halide:** When the carbon of an alkane is bonded to one or more halogens, the resulting compound is called an **alkyl halide** or **haloalkane**. For years, chloroform (Figure 8.6.2), a haloalkane with the formula CHCl_3 , was a commonly used solvent in the laboratory. This substance was also one of the earlier anesthetic drugs used in surgery. Its use is now highly restricted due to negative health effects, but it remains an important industrial chemical used in the production of PTFE (TeflonTM). Chlorodifluoromethane (see Figure 8.6.2) was used as a refrigerant and in aerosol sprays until the late twentieth century, but its use was discontinued after it was found to have harmful effects on the ozone layer. Bromoethane (see Figure 8.6.2) is a simple alkyl halide often used in organic synthesis. Alkyl halide groups are quite rare in biomolecules.

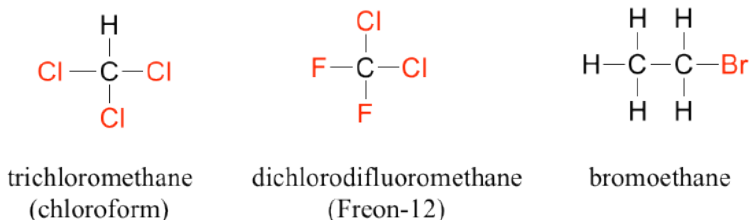


Figure 8.6.2 Structures of trichloromethane (chloroform), dichlorodifluoromethane (Freon 12), and bromoethane.

- **Alcohol:** In the **alcohol** functional group, a carbon is single-bonded to an **OH** group (the **OH** group, by itself, is referred to as a **hydroxyl**). Except for methanol, all alcohols can be classified as primary, secondary or tertiary:
 - A primary alcohol is an alcohol in which the carbon atom

carrying the hydroxyl group is bonded to only one other carbon atom.

- A secondary alcohol is an alcohol in which the carbon atom carrying the hydroxyl group is bonded to two other carbon atoms.
- A tertiary alcohol is an alcohol in which the carbon atom carrying the hydroxyl group is bonded to three other carbon atoms.

When the hydroxyl group is *directly* attached to an aromatic ring, the resulting group is called a **phenol**. The sulphur analog of an alcohol is called a **thiol** (from the Greek *thio*, for sulphur). The following figure displays the different types of alcohol that we discussed in this section:

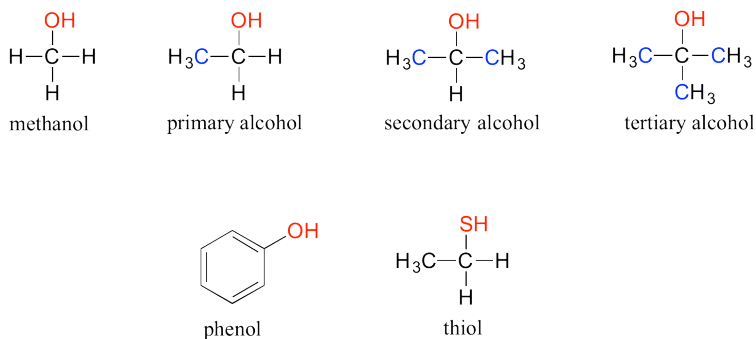
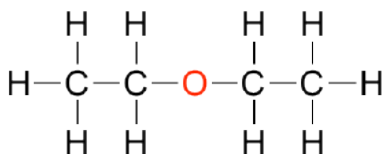
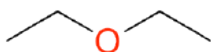


Figure 8.6.3 Different types of alcohols.

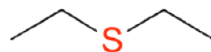
- **Ether:** In an **ether** functional group, a central oxygen is bonded to two carbons. Below is the structure of diethyl ether (see Figure 8.6.4), a common laboratory solvent and one of the first compounds to be used as an anaesthetic during operations. The sulphur analog of ether is called a **thioether** or **sulphide**.



||



an ether



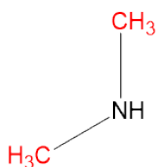
a sulfide

Figure 8.6.4 Different types of ether.

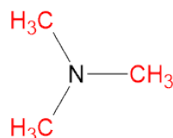
- Amines:** **Amines** are characterised by nitrogen atoms with single bonds to hydrogen and carbon. Just as there are primary, secondary, and tertiary alcohols, there are primary, secondary, and tertiary amines (see Figure 8.6). Ammonia is a special case with no carbon atoms.



Primary amine



Secondary amine



Tertiary amine

Figure 8.6.5 Amines

- Carbonyl groups:** A number of functional groups contain a carbon-oxygen double bond, commonly referred to as a **carbonyl**. **Ketones** and **aldehydes** are two closely related carbonyl-based functional groups that react in very similar

ways. In a ketone, the carbon atom of a carbonyl is bonded to two other carbons. In an aldehyde, the carbonyl carbon is bonded on one side to hydrogen and on the other side to carbon. The exception to this definition is formaldehyde, in which the carbonyl carbon has bonds to two hydrogens. Examples of aldehydes and ketones are illustrated in Figure 8.6.6.

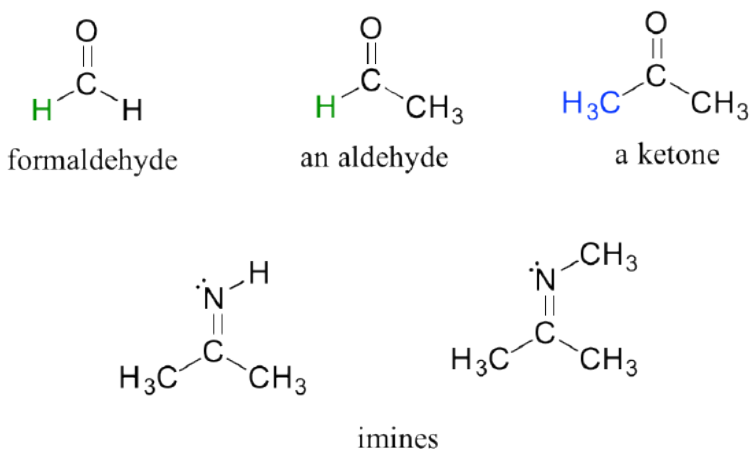


Figure 8.6.6 Aldehydes, ketones and imines.

Imines: An imine is a functional group in organic chemistry with a carbon-nitrogen double bond ($\text{C}=\text{N}$). It is derived from the reaction between a primary amine (RNH_2) and a carbonyl compound, typically an aldehyde or a ketone. Examples of imines are shown in Figure 8.6.6.

When a carbonyl carbon is bonded on one side to a carbon (or hydrogen) and on the other side to an oxygen, nitrogen, or sulphur, the functional group is considered to be one of the '**carboxylic acid derivatives**', a designation that describes a set of related functional groups. The eponymous member of this family is the **carboxylic acid** functional group, in

which the carbonyl is bonded to a hydroxyl group. Other derivatives are carboxylic esters (usually just called '**esters**') and **amides**. Other carboxylic acid derivatives also exist. Many are common in biology. Examples of carboxylic acids, esters, amides, acid chlorides and acid anhydrides are shown below (see Figure 8.6.7)

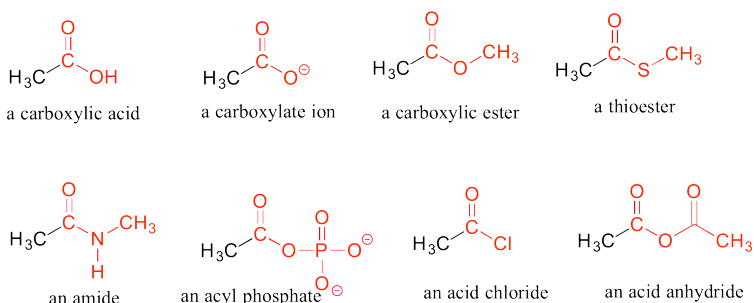
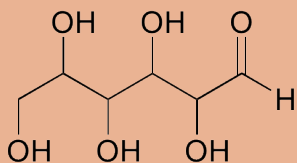


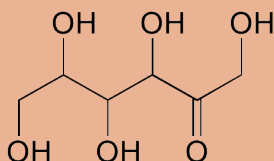
Figure 8.6.7 Examples of carboxylic acids, esters, amides, acid chlorides and acid anhydrides.

Example 8.6.1

A single compound often contains several functional groups, particularly in biological organic chemistry. The six-carbon sugar molecules glucose and fructose shown below contain aldehyde and ketone groups, respectively, and both contain five alcohol groups.

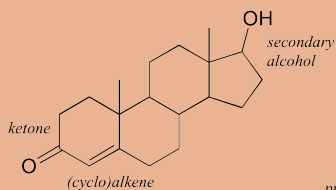


glucose

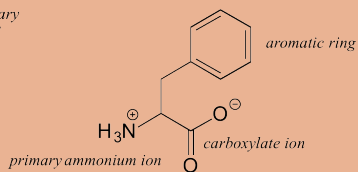


fructose

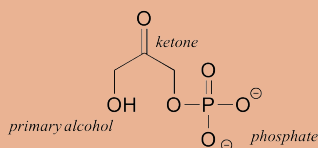
The hormone testosterone, the amino acid phenylalanine, and the glycolysis metabolite dihydroxyacetone phosphate all contain multiple functional groups, as labelled below.



testosterone



phenylalanine



dihydroxyacetone phosphate
(DHAP)

While not a complete list, this section has covered most of the

important functional groups we will encounter in biological organic chemistry.

Functional Groups and Organic Nomenclature

As noted earlier, the presence of a functional group frequently appears in the IUPAC name as a suffix. For alkanes, the names end in 'ane,' which indicates the absence of any functional group.

Alkenes are designated with an 'ene' ending. Compounds with multiple double bonds are called dienes, trienes, etc, as shown below in Figure 8.6.8.

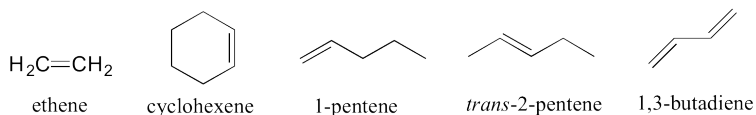


Figure 8.6.8 Nomenclature of double and triple bonds.

Some groups can only be present on a terminal carbon. Thus, a locating number is not necessary: aldehydes end in 'al', carboxylic acids in 'oic acid', and their conjugate base carboxylates in 'oate', as shown below in Figure 8.6.9.

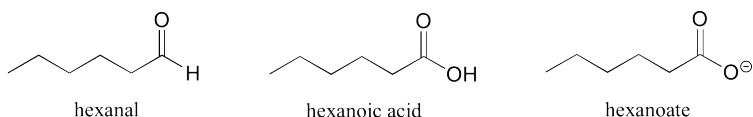


Figure 8.6.9 Nomenclature of carbonyl compounds.

Other functional groups have their suffixes, as well, and some functional groups affect IUPAC names in more complex ways. Many molecules also have multiple functional groups, complicating the names further.

It is not crucial to learn the details now, but it is valuable to

know that the suffix can often be used to identify the presence of a specific functional group on a molecule.

Key Takeaways

- The functional groups in organic molecules determine their reactivity.
- Some functional groups, such as carbonyl groups, amines, and alkyl halides, contain heteroatoms, meaning atoms other than carbon and hydrogen.
- Alkyl halides are alkanes with halogens.
- Alcohols are classified into primary, secondary and tertiary based on the arrangement of carbon atoms around the carbon bonded to the hydroxyl group.
- Carbonyl compounds such as aldehydes, ketones and carboxylic acids contain a carbon oxygen double bond.
- Complex organic molecules can contain more than one functional group.
- The functional group present in an organic molecule contributes to its IUPAC name.

Exercises

Practice Questions



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Transcript

8.7 Types of Organic Reactions

Learning Objectives

- Describe different reaction types in organic chemistry.
- Predict the products obtained upon reactions carried out by hydrocarbons, alkyl halides and alcohols.

The practice of organic chemistry includes understanding, planning for and carrying out functional group transformations of these types. A knowledgeable organic chemist can link together reactions to build specific target molecules according to a plan. Organic chemical synthesis has applications in many industries, from food to pharmaceuticals, adhesives and coatings, and more.

Main Reaction Types in Organic Chemistry

Organic chemistry involves a wide variety of reactions, each with its own set of conditions and mechanisms. Reaction types can be categorised based on the changes that occur in the structure of the molecules involved.

- Addition Reaction:** In addition reactions, atoms or groups of atoms are added, typically to a double or triple bond. In these reactions, one of the bonds in the double bond is broken. Each of the carbon atoms in the bond can then attach another atom or group while remaining joined to each other by a single bond. An example is the reaction between ethene and bromine. (See Figure 8.7.1). In this reaction, bromine atoms are added across the carbon-carbon double bond in ethene. The addition reaction with bromine can be used to test for alkenes in an uncharacterised sample of material. Bromine solutions are brownish-red. When we add a Br_2 solution to an alkene, the colour of the solution disappears because the alkene reacts with the molecular bromine.

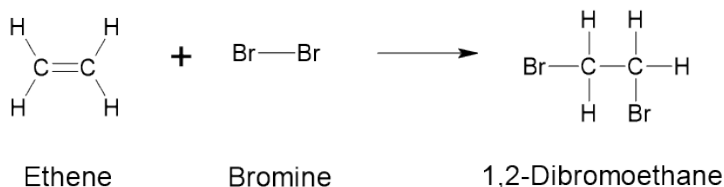
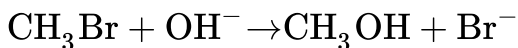


Figure 8.7.1 Example of an addition reaction.

- Substitution Reactions:** Substitution reactions involve the replacement of one functional group or atom by another. Let's consider the reaction between bromomethane (CH_3Br) and hydroxide ion (OH^-) as shown in the following equation. The bromine atom in bromomethane is replaced by the hydroxide ion, forming methanol and leaving the bromide ion.



- Elimination Reactions:** Elimination reactions involve the removal of elements from a molecule, leading to the formation

$$\begin{array}{c}
 \text{H} & \text{H} & \text{H} & \text{H} \\
 | & | & | & | \\
 \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{Br} \\
 | & | & | & | \\
 \text{H} & \text{H} & \text{H} & \text{H}
 \end{array}
 \xrightarrow{\text{OH}^- \text{ (base catalyst)}}
 \begin{array}{c}
 & & \text{H} \\
 & & | \\
 \text{H} & \text{H} & \text{C}-\text{H} \\
 | & | & / \backslash \\
 \text{H}-\text{C}-\text{C} & \text{C} & \text{H} \\
 | & | & | \\
 \text{H} & \text{H} & \text{H}
 \end{array}
 + \text{HBr}$$

1-Bromobutane But-1-ene Hydrogen bromide

- **Oxidation and Reduction Reactions:** As you learned in Chapter 4.6 (red-ox reactions), reduction is the gain of electrons, and oxidation is the loss of electrons of an element or atom. The same principle applies in organic chemistry as well. Nevertheless, due to the intricate nature of organic molecules, organic chemists use the following conventions to identify oxidation and reduction reactions:
 - **Oxidation** is when there is an increase in the number of $\text{C} - \text{O}$ bonds and or a decrease in the number of $\text{C} - \text{H}$ bonds. Figure 8.7.3 displays the oxidation of alcohol into aldehyde and further oxidation leading to the production of carboxylic acid. You can see that the number of $\text{C} - \text{O}$ bonds increase when ethanol converts to acetaldehyde and acetaldehyde converts to acetic acid.

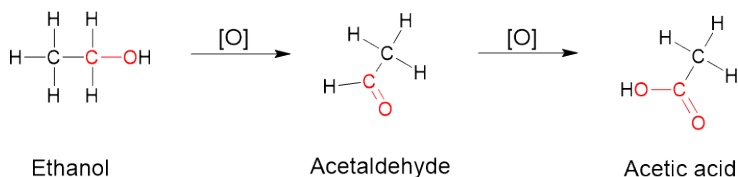


Figure 8.7.3 Example of an oxidation reaction.

- Reduction** is the opposite of oxidation. A decrease in the number of $\text{C} - \text{O}$ bonds and/or an increase in the number of $\text{C} - \text{H}$ bonds is considered a reduction. We often observe reduction reactions as the addition of hydrogen across a double or triple bond, as shown in Figure 8.7.4. Reduction reactions are a specific instance of an addition reaction.

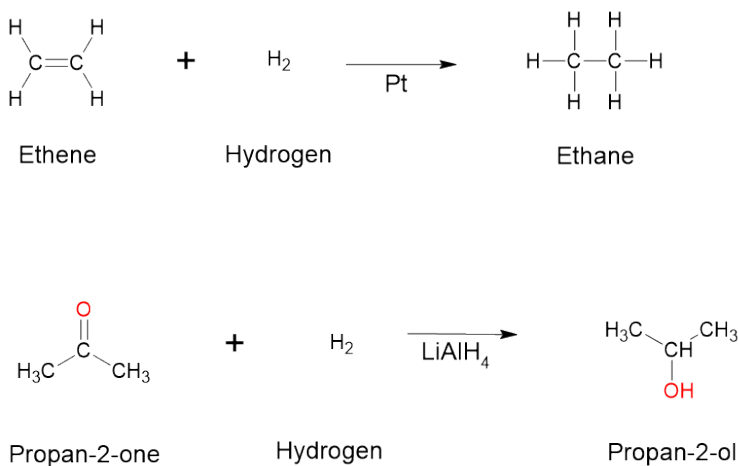


Figure 8.7.4 Examples of reduction reactions.

- Condensation Reactions:** Condensation reactions refer to a

class of reactions in which two molecules combine to form a larger molecule, often with the elimination of a smaller molecule, such as water or alcohol.

- One of the well-known condensation reactions is the esterification reaction between a carboxylic acid and an alcohol. In this reaction, a water molecule is eliminated, and an ester is formed. For example, the esterification of acetic acid with ethanol results in the formation of ethyl acetate and water, as illustrated in Figure 8.7.5.

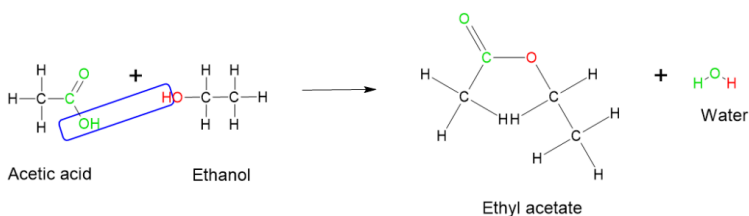


Figure 8.7.5 Esterification reaction.

- Condensation reactions also occur when amines react with carboxylic acids or acid chlorides, producing an amide (see Figure 8.7.6). In the example below, acetic acid reacts with ammonia to produce acetamide, leaving a water molecule. In a protein, individual amino acids are joined together by amide bonds. The amide bond, also known as a peptide bond, is formed through a condensation reaction between the carboxyl group of one amino acid and the amino group of another amino acid. This reaction involves the elimination of a water molecule.

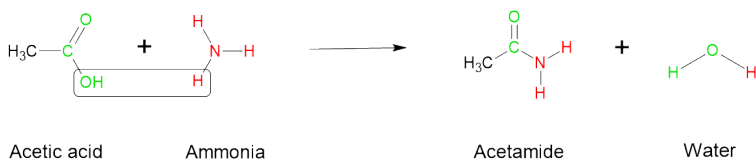


Figure 8.7.7 Amide formation.

- Hydrolysis Reactions:** Hydrolysis reactions involve the cleavage of a bond with the addition of water. This process often results in the breakdown of larger molecules into smaller components. Esters and amides can be reversed into their original reactants by a hydrolysis reaction, as shown in Figure 8.7.8. Hydrolysis is the opposite of the condensation reactions.

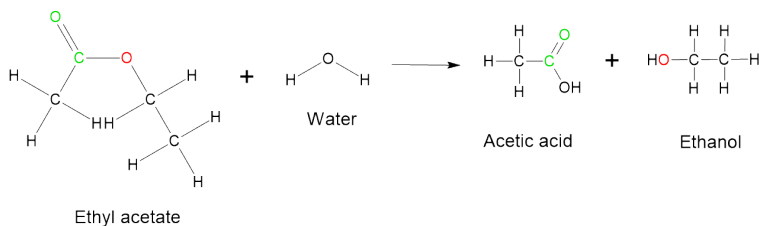
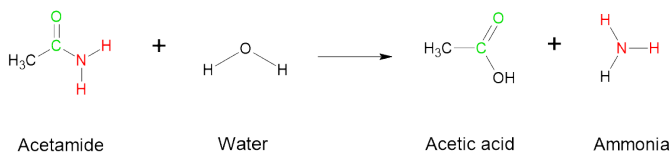


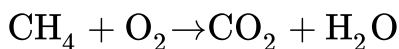
Figure 8.7.8 Hydrolysis reactions.

Reactions of Hydrocarbons

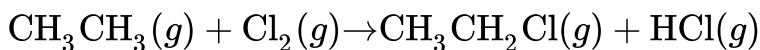
Reactions of alkanes

As we learned in Chapter 8.1, alkanes are relatively unreactive compared to other classes of organic compounds due to the non-polar nature and strength of carbon-carbon and carbon-hydrogen single bonds. The only major reactions that alkanes undergo are combustion and substitution reactions.

- **Combustion reactions:** in combustion reactions, alkanes react with oxygen to produce carbon dioxide and water, as shown in the following chemical equation for methane. This reaction releases a significant amount of energy and is the basis for the use of hydrocarbons as fuels. Methane is the main natural gas used in home heating and cooking.



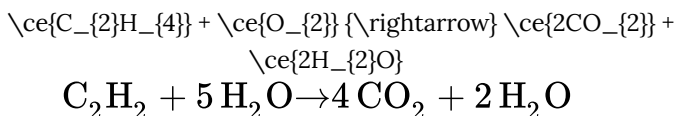
- **Substitution reactions:** substitution reactions of alkanes involve the replacement of one or more hydrogen atoms in an alkane with other atoms or groups. The most common type of substitution reaction for alkanes is halogenation, where hydrogen atoms are replaced by **halogen** atoms (chlorine, bromine, etc.). This process typically occurs under conditions of high temperature or ultraviolet (UV) light. Chloromethane, which is a general anaesthetic, can be produced by the reaction between chlorine and methane under high temperatures, as displayed in the following equation.



Reactions of alkenes and alkynes

As alkenes and alkynes are unsaturated hydrocarbons, they are more reactive than alkanes.

Alkenes and alkynes also undergo **combustion reactions** to form carbon dioxide and water as alkanes:



Except for combustion reactions, nearly all other reactions of alkenes and alkynes often involve the breaking of double or triple bonds to form new bonds by addition reactions. The different types of addition reactions of alkenes and alkynes are discussed below.

- **Hydrogenation (addition of hydrogen):** The simplest addition reaction is **hydrogenation**. This is a reaction in which hydrogen gas reacts at a carbon-to-carbon double or triple bond to add hydrogen atoms to carbon atoms. In the laboratory, this reaction can be facilitated with hydrogen in the presence of a catalyst such as nickel (**Ni**) or platinum (**Pt**). Examples of hydrogenation reactions are displayed in Figure 8.7.9. The hydrogenation reaction is a specific type of addition reaction, which is also a reduction reaction, as discussed in reduction reactions earlier in this section. Hydrogenation is used to convert unsaturated vegetable oils to saturated fats. Hydrogenated fats, or partially-hydrogenated fats, often appear on nutrition labels for processed foods such as crackers or chips.

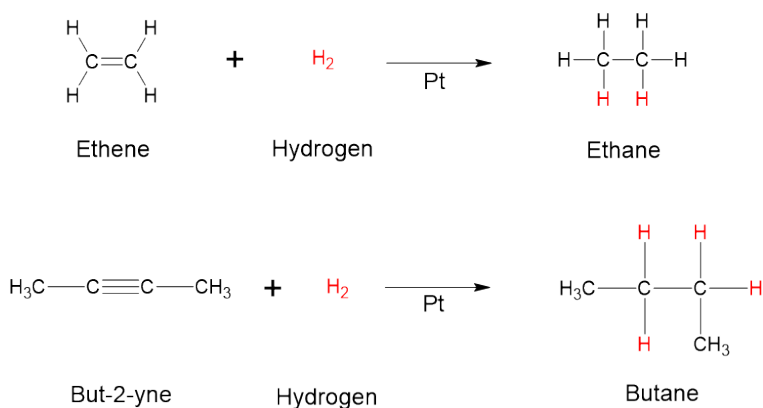


Figure 8.7.9 Hydrogenation of alkenes and alkynes.

- Halogenation (addition of halogens such as chlorine and bromine):** alkenes and alkynes also readily undergo halogenation through addition, a reaction in which a halogen (chlorine or bromine) reacts at a carbon-to-carbon double or triple bond to add halogen atoms to carbon atoms. The reaction involves the addition of one halogen atom to each carbon atom of the double bond or two halogen atoms to each carbon atom of a triple bond (Figure 8.7.10).

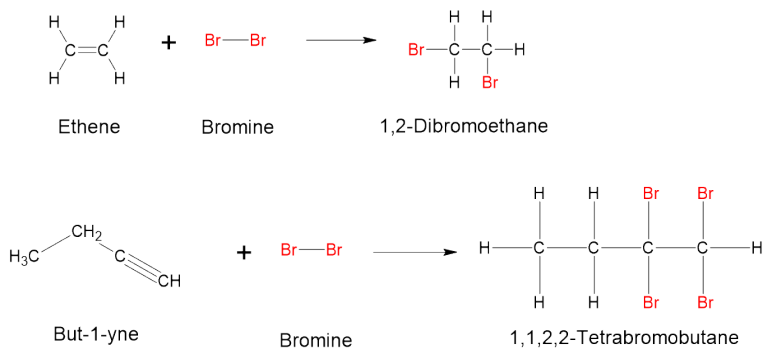


Figure 8.7.10 Halogenation of alkenes and alkynes

- **Hydrohalogenation (Addition of hydrogen halides HX):** Alkenes and alkynes react with hydrogen halides such as **HCl**, and **HBr** to produce alkyl halides (**RX**). The hydrogen and the halogen atoms are added across the double or triple bond. The following equation displays the addition of hydrogen chloride to ethene.



- **Hydration (addition of water):** Another important addition reaction occurs between an alkene and water to form an alcohol. This reaction is called hydration and represents the addition of water to a substance. In this reaction, water is added across the double bond in the presence of an acid catalyst. The following figure displays an example of a hydration reaction of an alkene.

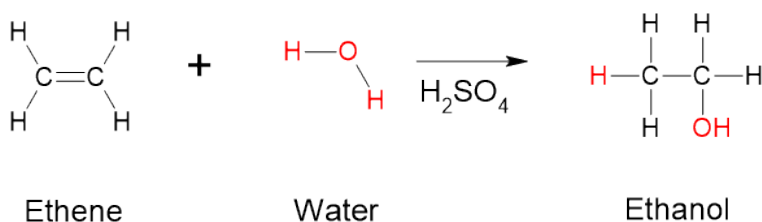


Figure 8.7.11 Hydration reaction.

Example 8.7.1

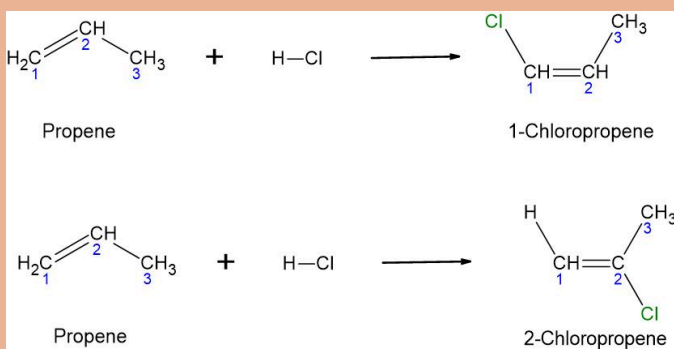
Problem

What are the possible products of the addition of the following reagents to propene $\text{CH}_2=\text{CHCH}_3$:

1. HCl
2. H_2O

Solution

1. The addition of HCl is a hydrohalogenation. One carbon atom of the double bond receives hydrogen, and the other carbon in the double bond receives the chlorine atom. Based on this definition, two products are possible, as shown in the following figure. In the first product, the chlorine atom is bound to the first carbon, and in the second product, the chlorine atom is bound to the second carbon.

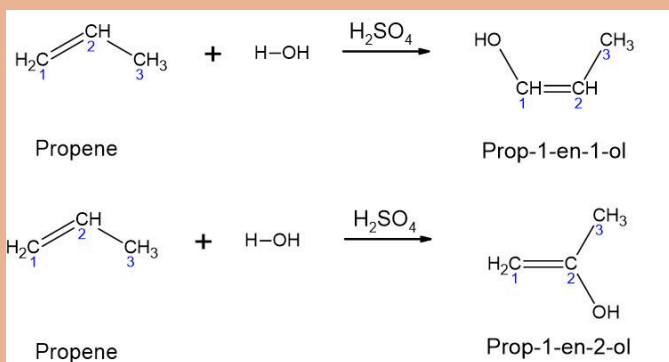


So, how do we decide which product is actually formed? When unsymmetrical alkenes undergo hydrohalogenation and hydration reactions, more than one product is possible. **Unsymmetrical alkenes** means an alkene in which the two carbon atoms of the double bond are not equivalently substituted. For instance, in propene, the first carbon of the double bond has two hydrogens, and the second carbon of the double bond has one hydrogen and one methyl group.

Therefore, in the above reaction, both products are possible to form in reality. In such situations, one product is generally dominant over the other. The dominant product is called the 'major product' of the reaction. The major product can be predicted by Markovnikov's rule. You will learn how to use Markovnikov's rule to predict products of unsymmetrical alkenes in your university organic chemistry course. As this book is an introductory chemistry book, Markovnikov's rule will not be

covered in this book. If you are interested to know more about Markovnikov's rule, visit this webpage.

2. The same concept as above applies to the hydration of alkenes. The possible products are $\text{CH}(\text{OH})=\text{CH}-\text{CH}_3$ and $\text{CH}_2=\text{C}(\text{OH})-\text{CH}_3$, as shown in the following figure.



Addition polymerisation of alkenes

Addition polymerisation is a process in which monomers with unsaturated double bonds react to form a polymer without the elimination of any byproducts. The reaction involves the successive

addition of monomers to the growing polymer chain. Addition polymerisation is characteristic of alkenes and results in the formation of polymers with saturated carbon backbones. Common examples of addition polymerisation include the polymerisation of ethene (ethylene) to form polyethylene (Figure 8.7.12) and the polymerisation of propene (propylene) to form polypropylene. The resulting polymers have high molecular weights and find extensive use in various applications due to their physical and chemical properties.

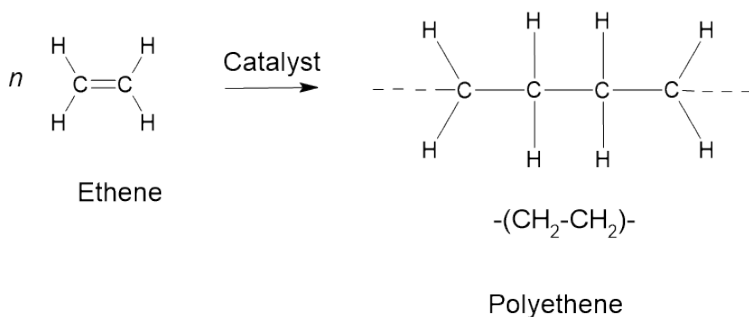


Figure 8.7.12 Addition Polymerisation

Reactions of functional groups

Reactions of alkyl halides

Alkyl halides, also known as haloalkanes, are compounds in which one or more hydrogen atoms in an alkane have been replaced by halogen atoms (fluorine, chlorine, bromine, or iodine). Alkyl halides differ in polarity compared to alkanes due to the presence of a

halogen atom. In alkyl halides, a halogen atom is attached to a carbon atom in place of a hydrogen atom. Halogens are more **electronegative** than carbon, leading to a significant electronegativity difference. This electronegativity difference results in a polar covalent bond between carbon and the halogen. Therefore, the halogen end of the molecule carries a partial negative charge (δ^-), and the carbon end carries a partial positive charge (δ^+). Hence, the carbon atom with a partial positive charge is attacked by the negative ions. As a result, alkyl halides undergo **substitution reactions**, as shown below in Figure 8.7.13. For instance, the chlorine atom in chloromethane is replaced by a hydroxyl group, forming methanol. Alcohols can be prepared by using reactions between alkyl halides and a base. Alkyl halides undergo substitution reactions with many groups, such as ammonia, amines, cyanide, hydrosulphide and water.

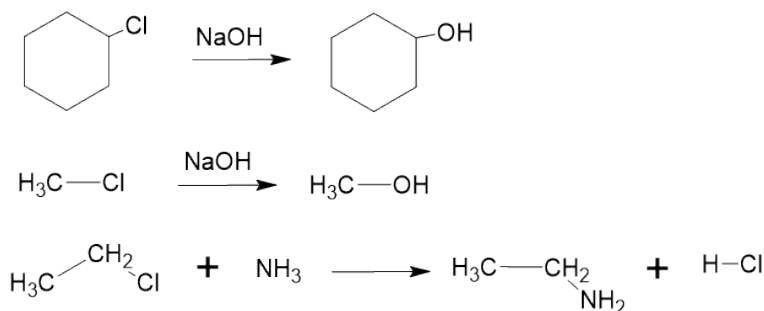


Figure 8.7.13 Substitution reactions of alkyl halides.

Reactions of alcohols

Alcohols mainly undergo substitution and oxidation reactions.

- Substitution reactions: alcohols can be converted to alkyl halides (alkyl bromides, alkyl chlorides, etc.) through

substitution reactions. Primary and secondary alcohols are converted to alkyl halides using thionyl chloride SOCl_2 , phosphorus tribromide PBr_3 or phosphorus trichloride PCl_3 . Tertiary alcohols are converted into alkyl halides using HCl or HBr .

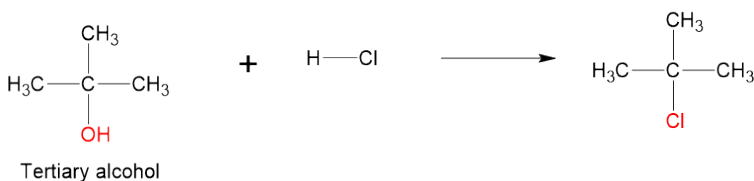
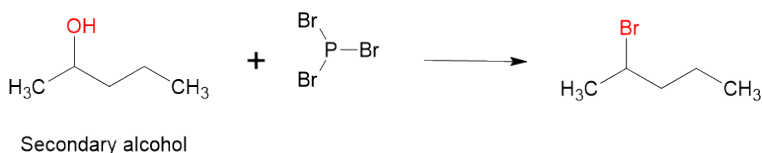
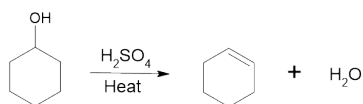


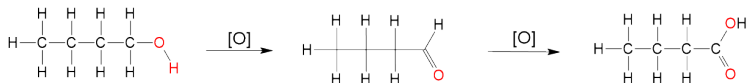
Figure 8.7.14 Substitution reactions of alcohols.

- Oxidation reactions (Figure 8.7.15):
 - Primary alcohols can be oxidised to form aldehydes and then further oxidised to carboxylic acids. Mild oxidising agents such as pyridinium chlorochromate are used to convert a primary alcohol to an aldehyde. Common oxidising agents such as potassium permanganate KMnO_4 and potassium dichromate $\text{K}_2\text{Cr}_2\text{O}_7$ convert primary alcohol to carboxylic acid.
 - Secondary alcohols can be oxidised to form ketones. Further oxidation is not generally favourable.
 - Tertiary alcohols are resistant to oxidation under normal conditions.
- Dehydration (Figure 8.7.15): alcohols can undergo dehydration to form alkenes in the presence of heat and an acid catalyst.

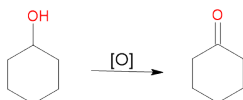
Dehydration



Oxidation



Primary alcohol



Secondary alcohol

Figure 8.7.15 Dehydration and oxidation reactions of alcohols.

Key Takeaways

- Organic chemistry involves comprehending, planning, and executing functional group transformations, allowing skilled chemists to link reactions and construct targeted molecules systematically.
- Major reaction types in organic chemistry include addition, elimination, substitution, oxidation, reduction, condensation and hydrolysis.
- Due to the less reactivity of alkanes, they only undergo combustion and substitution reactions.

- Alkenes and alkynes undergo different types of addition reactions, including hydrogenation, halogenation and hydrohalogenation.
- The polarity of alkyl halides makes them undergo substitution reactions.
- Alcohols primarily undergo substitution and oxidation reactions.

Exercises

Practice Questions



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=2502#h5p-79>



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=2502#h5p-92>

Predict the products of the following oxidation reactions of alcohols with oxidising agents.



An interactive H5P element has been excluded from this version of the text. You can view it online here:

<https://rmit.pressbooks.pub/rmitchemistrybridgingcourse/?p=2502#h5p-93>

Transcript

PART IX

TRANSCRIPTS

2. 1.1 States of Matter

Practice questions

1. Which of the following is not an example of matter?
 - a. Hate
 - b. A tree
 - c. A book
 - d. A fried egg
2. Which of the following is a chemical property?
 - a. Steel wool burns when ignited by a flame.
 - b. A gallon of milk weighs over eight pounds.
 - c. Gold can be stretched into very thin wires.
 - d. Sulphur is yellow.
3. Which of the following statements represents a physical change? *Select all possible answers.*
 - a. The alcohol in many thermometers freezes at about -40 degrees Fahrenheit.
 - b. The elements sodium and chlorine come together to make a new substance called sodium chloride.
 - c. Water boils and becomes steam.
 - d. Food is converted into usable form by the digestive system.
4. Which of the following is an element?
 - a. Carbon
 - b. Water
 - c. Coffee
 - d. Air
5. Which of the following is a pure substance?
 - a. Hydrogen
 - b. Seawater

- c. Tea
 - d. Concrete
6. Which of the following is a heterogeneous mixture?
- a. Chocolate chip cookie
 - b. Carbon
 - c. An emulsion
 - d. Salt
7. Which of the following statements is true?
- a. All heterogeneous mixtures must contain three or more components.
 - b. Pure substances cannot be broken down into simpler substances by physical methods.
 - c. Sugar is an element.
 - d. A mixture of water, salt and sand is an example of a homogeneous mixture.
8. Which of the following substances takes the shape of the container and has a definite volume?
- a. Stainless steel wire
 - b. Salt crystals
 - c. Helium gas
 - d. A glass of fruit juice
9. Which statement correctly describes the volume and shape of a gas?
- a. It has an indefinite shape and a fixed volume.
 - b. It has a fixed shape and an indefinite volume.
 - c. It has a fixed shape and a fixed volume.
 - d. It takes the shape of the container and expands to fill the container.
10. Which of the following statements describes a chemical property? *Select all the correct answers.*
- a. Copper objects turn green when exposed to moist air for a prolonged period.
 - b. The density of Lithium metal is 0.534g/ml at 20°C.

- c. Potassium metal reacts vigorously with water to form potassium hydroxide solution and hydrogen gas.
 - d. Aluminium metal is a solid at room temperature.
11. Why is iron considered a metal?
- a. It is a solid at room temperature.
 - b. It is shiny.
 - c. It conducts electricity and heat well.
 - d. All of the above.
12. Fill in the blanks with correct words from the following list:
phases, colours, electricity, heat, ductile, shiny, brittle, heat, malleable, do not, electricity
- Begin activity: Metals are typically
conduct..... andwell, and
are..... and..... . Whereas non-metals are a variety
of and , are in the solid phase,
and conduct or well.
13. Fill in the blanks with correct words from the following list:
brittle, heat, metal, non-metal, sheets, wires, electricity
- Begin activity:
- Carbon behaves as a because it conducts
and well. It is a because it is black and
..... and cannot be made into or

Solutions

- 1. (a)
- 2. (a)
- 3. (a) and (c)
- 4. (a)
- 5. (a)
- 6. (a)
- 7. (b)

8. (d)
9. (d)
10. (a) and (c)
11. (d)
12. Metals are typically **shiny**, conduct **electricity** and **heat** well, and are **malleable** and **ductile**. Whereas non-metals are a variety of **colours** and **phases**, are **brittle** in the solid phase, and **do not** conduct **heat** or **electricity** well.
13. Carbon behaves as a **metal** because it conducts **heat** and **electricity** well. It is a **non-metal** because it is black and **brittle** and cannot be made into **sheets** or **wires**.

3. 1.2 Expressing Numbers

Practice questions

1. Express these numbers in scientific notation.
 - a. **56.9**
 - b. **563, 000**
 - c. **0.0804**
 - d. **0.00000667**
2. Express these numbers in scientific notation.
 - a. **0.00656**
 - b. **65, 600**
 - c. **4567, 000**
 - d. **0.000005507**
3. Express these numbers in standard notation.
 - a. **1.381×10^5**
 - b. **5.22×10^{-7}**
 - c. **9.998×10^4**
 - d. **8.09×10^0**
 - e. **3.088×10^{-5}**
4. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.
 - a. **72.44×10^3**
 - b. **$9, 943 \times 10^{-5}$**
 - c. **$588, 399 \times 10^2$**
 - d. **0.234×10^{-3}**
5. How many significant figures do these numbers have?
 - a. **23**

- b. 23.0
 - c. 0.00023
 - d. 0.0002302
 - e. 1.2000×10^5
6. Compute and express each answer with the proper number of significant figures, rounding as necessary.
- a. $56.0 + 3.44 = ?$
 - b. $0.00665 + 1.004 = ?$
 - c. $45.99 - 32.8 = ?$
 - d. $45.99 - 32.8 + 75.02 = ?$
7. Compute and express each answer with the proper number of significant figures, rounding as necessary.
- a. $56.7 \times 66.99 = ?$
 - b. $\frac{1000}{77} = ?$
 - c. $\frac{1000}{77.0} = ?$
 - d. $6.022 \times 1.89 = ?$

Solutions

1. Scientific notations are as below.

- a. 5.69×10^1
- b. 5.631×10^5
- c. 8.04×10^{-2}
- d. 6.67×10^{-6}

2. Scientific notations are as below.

- a. 6.56×10^{-3}
- b. 6.56×10^4
- c. 4.567×10^6

- d. 5.507×10^{-6}
3. The answers are as follows.
- a. 138,100
 - b. 0.000000522
 - c. 99,980
 - d. 8.08
 - e. 0.00003088
4. The answers are as follows.
- a. 7.244×10^4
 - b. 9.943×10^{-2}
 - c. 5.88399×10^7
 - d. 2.34×10^{-4}
5. The answers are as follows.
- a. 2
 - b. 3
 - c. 2
 - d. 4
 - e. 5
6. The answers are as follows.
- a. 59.4
 - b. 1.011
 - c. 13.2
 - d. 88.2
7. The answers are as follows.
- a. 3.80×10^3
 - b. 0.013
 - c. 0.0130
 - d. 11.4

4. 1.3 Units and Conversions

Practice questions

- Identify the unit in each quantity.
 - 2 boxes of crayons
 - 3.5 grams of gold
 - 32 oz of cheddar cheese
 - 0.045cm^3 of water
 - 2 dozen eggs
 - 2.4km/s
- Indicate what multiplier each prefix represents. Write the power of 10 in the space next to x .
 - $k - 10^x, x = ?$
 - $m - 10^x, x = ?$
 - $c - 10^x, x = ?$
 - $G - 10^x, x = ?$
 - $\mu - 10^x, x = ?$
- Give the prefix that represents each multiplier:
 - $\frac{1}{1000} \times$
 - $1000 \times$
 - $1000000000 \times$
 - $\frac{1}{1000000000} \times$
- Express each quantity in a more appropriate unit. Choose the correct answer.
 - $0.00000344\text{s} - 3.44\mu\text{s}, 3.44\text{ns}, 3.44\text{Ms}$
 - $3500\text{L} - 3.35\text{kL}, 3.35\text{mL}, 3.35\text{GL}$

- c. $0.045ms - 4.5cm, 4.5mm, 4.5km$
5. Write the two conversion factors that exist between the two given units.
- millilitres and litres =
 - microseconds and seconds =
 - kilometres and metres =
6. Perform the following conversions and write the answer.
- $2.56gal$ gallons to litres ($1L = 0.265gal$) =
 - $750.0\mu mol$ to millimoles =
 - $45km/h$ to m/s =
 - $5.4km$ to metres =
 - $0.665m$ to millimetres =

Answers

- boxes
 - grams
 - oz
 - cm^3
 - dozen
 - km/s
- 3
 - 3
 - 6
 - 2
 - 9
 - 6
- Mili
 - kilo
 - Giga
 - Nano

a. $3.44\mu s$

b. $3.35kL$

c. $4.5cm$

a. $\frac{1000mL}{1L}$ and $\frac{1L}{1000mL}$

b. $\frac{1000000\mu s}{1s}$ and $\frac{1s}{1000000\mu s}$

c. $\frac{1000m}{1km}$ and $\frac{1km}{1000m}$

a. $10.0L$

b. $0.7500mmol$

c. $12.5m/s$

d. $5400m$

e. $665mm$

5. 1.4 Temperature and Density

1. Perform the following conversions:
 - a. $255^{\circ}F$ to degrees Celsius
 - b. $-255^{\circ}F$ to degrees Celsius
 - c. $50.0^{\circ}C$ to degrees Fahrenheit
 - d. $-50.0^{\circ}C$ to degrees Fahrenheit
2. Perform the following conversions:
 - a. $100.0^{\circ}C$ to kelvins
 - b. $-100.0^{\circ}C$ to kelvins
 - c. $100K$ to degrees Celsius
 - d. $300K$ to degrees Celsius
3. Which of the following are possible units for density? *Select all the possible answers.*
 - a. g/mL
 - b. mL/km
 - c. g/L
 - d. kg/L
4. A sample of iron has a volume of $48.2cm^3$. What is its mass? Use the density table provided in the text to get the density of iron.
 - a. $379g$
 - b. $45.9g$
 - c. $678g$
 - d. $3.79g$
5. The volume of hydrogen used by the Hindenburg, the German airship that exploded in New Jersey in 1937, was $2.000 \times 10^8 L$. If hydrogen gas has a density of

$0.0899\text{g}/L$, what mass of hydrogen was used by the airship?

- a. $1.80 \times 10^7\text{g}$
 - b. $3.20 \times 10^4\text{g}$
 - c. $25.1 \times 10^7\text{g}$
 - d. $0.80 \times 10^7\text{g}$
6. What is the volume of 100.0g of lead if lead has a density of $11.34\text{g}/\text{cm}^3$?
- a. 10.54cm^3
 - b. 3.98cm^3
 - c. 6.818cm^3
 - d. 8.818cm^3
7. What is the volume in litres of 222g of neon if neon has a density of $0.900\text{g}/L$?
- a. $2.89L$
 - b. $563L$
 - c. $87.8L$
 - d. $247L$

Answers

- 1.
 - a. 124^0C
 - b. -159^0C
 - c. 122^0F
 - d. -58^0F
- 2.
 - a. $373K$
 - b. $173K$
 - c. -173^0C

d. 27°C

3. a, c, d

4. a

5. a

6. d

7. d

6. 2.1 Atoms: The Building Blocks of Everything

Practice questions

Multiple choice questions (There is/ are one or more correct answer(s) for each question).

1. Which of the following are statements of modern atomic theory?
 - a. All matter is composed of atoms.
 - b. All matter consists of energy.
 - c. Atoms of the same element are the same; atoms of different elements are different.
 - d. Atoms combine in whole-number ratios to form compounds.
 - e. Energy and matter can be transformed into one another.
2. Which subatomic particle is larger?
 - a. A proton
 - b. An electron
3. Which subatomic particle is larger?
 - a. A neutron
 - b. An electron
4. Where is most of the mass of an atom located?
 - a. Protons
 - b. Electrons
 - c. Neutron
 - d. Nucleus

Short-answer questions

1. What are the charges for each of the three subatomic particles?
 - a. Protons
 - b. Electrons
 - c. Neutrons
2. Calculate the number of protons, electrons and neutrons present in the atom Argon, whose atomic number is 18 and the mass number is 40.
3. The atomic number of Lithium is 3, and its mass number is 7. Calculate the total number of nucleons present in the atom.

Solutions

Multiple choice questions (There are one or more correct answer(s) for each question)

1. a, c, d
2. a
3. a
4. d

Short-answer questions

- a. Protons – Positive
- b. Neutrons – no charge
- c. Electrons – Negative
- a. Number of protons=Atomic number Z =Number of

electrons = 18

- b. Number of neutrons = Mass number - Number of protons
- c. Number of neutrons = $40 - 18$
- d. Number of neutrons = 22
- a. Protons and neutrons are nucleons.
- b. Total number of nucleons = Number of protons + Number of neutrons
- c. Mass number = Number of protons + Number of neutrons
- d. Therefore, the total number of nucleons is equal to the Mass number, which is 40.

7. 2.2 Electronic Configuration

Figures

Figure 2.2.4: Boron Electronic Diagram.

An electronic energy diagram for a boron atom. A vertical arrow on the left denotes increasing energy levels. Subshells 1s, 2s, 2p, 3s and 3p are written from bottom to top, with horizontal lines representing electron orbitals: s subshells have a single horizontal line, while p subshells have 3 lines in each.

An up and a down arrow are found on both the 1s and 2s horizontal line, representing 4 electrons in total. The 2p subshell has only 1 up arrow on the left horizontal line, representing an unpaired electron. There are 5 arrows in total, representing the 5 electrons needed to be allocated. The higher energy levels have no arrows and remain unfilled.

Figure 2.2.5: Carbon, Nitrogen and Oxygen Electronic Energy Diagrams

Electronic energy diagrams for Carbon, Nitrogen and Oxygen atoms, respectively. A vertical arrow on the left denotes increasing energy levels. Subshells 1s, 2s, 2p, 3s and 3p are written from bottom to top, with horizontal lines representing electron orbitals: s subshells have a single line, while p subshells have 3 lines in each.

On the left is Carbon. Both the 1s and 2s subshells have two arrows on each line, one pointing up and the other down, representing four electrons in total. The 2p subshell has two unpaired electrons, each represented by an up arrow on the first two underscores. The carbon atom has a total of six electrons, with the 3s and 3p energy levels remaining unfilled

Next, in the middle, is Nitrogen. The arrangement of electrons in the subshells is similar to Carbon, except there is an extra unpaired electron in the 2p subshell on the third horizontal line.

On the right is Oxygen. The arrangement of the electrons is nearly the same as Nitrogen, except there is now a paired electron in the first line of the 2p subshell.

Practice questions

Multiple choice questions

Question 1

1. Write the electron configuration of Na.

- a. $1s^2 2s^2 2p^5 3s^2$
- b. $1s^1 2s^2 2p^3 3s^4$
- c. $1s^2 2s^2 2p^6 3s^1$
- d. $1s^2 2s^2 2p^4 3s^2$

Question 2

1. Write the electron configuration of Ca using noble gas configuration.
 - a. $[Ar]3d^2$
 - b. $[Ca]3d^2$
 - c. $[Ar]4s^2$
 - d. $[Ca]4s^2$

Solutions

1. The first two electrons go to the 1s orbital as a pair. The next two electrons are placed in the orbital 2s. The next six electrons are placed in 2p orbitals as three pairs. The remaining electron is placed in the 3s orbital.
2. For Ca, Ar is the last noble gas that comes before Ca. The electron configuration of Ca is written as [Ar], which denotes Ar's electron configuration followed by the configuration of remaining electrons.

$Z = 18 : 1s^2 2s^2 2p^6 3s^2 3p^6$ (This is written as [Ar])
 $Z = 20 : 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ (After 3p orbital, electrons fill into the 4s orbital, not the 3d orbital). The noble gas configuration of Ca is $[Ar]4s^2$.

8. 2.3 The Periodic Table

Practice questions

Fill in the answer.

1. How many valence electrons does each of the following atoms have? (Hint: The group numbers of these elements can be used to determine the number of valence electrons. For groups 1–10, the group number equals the number of valence electrons. For groups 12–18, subtract 10 out of the group number to get the valence electrons. For group 11, the number of valence electrons can be anticipated as 11 due to the $s1$ electron configuration. However, there are a few exceptions in this case.)
 - a. P =
 - b. Sc =
 - c. F =
 - d. Ca =
 - e. Fe =

Solutions

- a. P = 2
- b. Sc = 3
- c. F = 7
- d. Ca = 2
- e. Fe = 8

9. 2.4 The Periodic Trends

Figures

Figure 2.4.3: Ionisation Energy on the Periodic Table.

A diagram of ionisation energies across the s and p blocks of the periodic table, showing varying sizes of the atoms with spheres varying in size, corresponding with ionisation values. The spheres represent atoms, with helium in the top right corner with the smallest sphere size and highest ionization energy of 2372 kJ/mol. Sizes increase as ionisation energy decreases towards caesium at the bottom left of the periodic table, with an ionisation value of 376 kJ/mol. Notable exceptions are found within group 16, being Te and Po.

Figure 2.4.4: Electronegativity of elements

A diagram of s and p blocks of the periodic table displaying the electronegativity values of the elements. The diagram has two large arrows indicating increasing electronegativity in the periodic table—the first arrow begins at the top left and goes to the right; and the second arrow starts from the bottom left and goes vertical to the top of the table. The electronegativity values decrease from fluorine in group 7A which has the highest electronegativity at 3.98; moves down to francium in group 1 with the lowest at 0.7. Noble gases do not have an electronegativity value.

Practice questions

Multiple choice questions

1. Predict which atom is larger for the following pair of atoms.
 - a. C
 - b. F
2. Predict which atom is larger for the following pair of atoms.
 - a. Na
 - b. Cs
3. In Figure 2.4.1, which element does not follow the standard atomic radius rule?
 - a. I
 - b. Te
 - c. Xe
 - d. Sb
4. Which of the following molecules has the greatest electronegativity?
 - a. Na
 - b. Cl
 - c. B
 - d. Ge

True/false question

1. Sodium has a greater atomic radius than chlorine; therefore, sodium has a higher ionisation energy.

Solutions

Multiple choice questions

1. (a).
C $Z=6$ and F $Z=9$ are located in the second period. When moving across a period from left to right, atomic radii decrease as a result of increasing Z . C lies to the left of F in the same row. Therefore, C is larger than F.
2. (b).
Na $Z=11$ and Cs $Z=55$ are located in group 1. Atomic radii increase from top to bottom of the periodic table as n increases. Cs is located below Na in the same column. Thus, Cs has a larger radius than Na.
3. (b).
It has a higher atomic radius than Sb, which is further right on the table.
4. (b).

True/false question

1. False

10. 2.5 Isotopes and Relative Abundance

Practice questions

Multiple choice questions

Select the option that best defines 'atomic mass unit' and that gives its correct abbreviation?

- a. The atomic mass unit is defined as one-twelfth of the mass of a carbon – 14 atom. Its abbreviation is AM
- b. The atomic mass unit is defined as one-twelfth of the mass of a carbon – 12 atom. Its abbreviation is amu.
- c. The atomic mass unit is defined as one-twelfth of the mass of a carbon – 12 atom. Its abbreviation is AM.
- d. The atomic mass unit is defined as one-sixth of the mass of a carbon – 12 atom. Its abbreviation is amu.

Short-answer questions

- 1. Estimate the mass, in whole numbers, of each isotope.
 - a. hydrogen
 - b. boron
 - c. iron
- 2. Estimate the mass, in whole numbers, of each isotope.
 - a. phosphorus
 - b. carbon

- c. americium
3. Determine the atomic mass of each element, given the isotopic composition. Provide answers to 3 decimal places.
- Lithium, which is 92.4 % lithium-7 (mass 7.016 amu) and 7.60% lithium-6 (mass 6.015 amu). What is the atomic mass of lithium?
 - Oxygen, which is 99.76 % oxygen-16 (mass 15.995 amu), 0.038% oxygen-17 (mass 16.999 amu), and 0.205% oxygen-18 (mass 17.999 amu). What is the atomic mass of oxygen?
4. Determine the atomic mass of each element, given the isotopic composition. Provide answers to 3 decimal places
- Neon, which is 90.48% neon-20 (mass 19.992 u), 0.27% neon-21 (mass 20.994 u), and 9.25% neon-22 (mass 21.991 u). What is the atomic mass of neon?
 - Uranium, which is 99.27 % uranium-238 (mass 238.051 u) and 0.720% uranium-235 (mass 235.044 u). What is the atomic mass of uranium?

Solutions

Multiple choice questions

1. (b)

Short-answer questions

- 1
 - 10
 - 56

2.
 - a. 31
 - b. 14
 - c. 241
3.
 - a. The atomic mass of lithium=6.940 amu/ 6.940 Da
 - b. The atomic mass of oxygen=16.000 amu/ 16.000 Da
4.
 - a. The atomic mass of neon= 20.180 u
 - b. The atomic mass of uranium=238.006 u

II. 3.1 The Octet Rule and Lewis Dot Diagrams

Figures

Figure 3.1.1: Lewis Electron Dot Diagrams

A diagram showing the Lewis electron-dot diagrams for neutral atoms, arranged by groups on the periodic table, with the first three rows of the table are shown. The dots represent electrons and may appear on the four sides of each elemental symbol with the number of dots representing the number of valence electrons. Note that it does not matter where the first dot is placed, and you must fill up each side with a single dot before pairing dots together when drawing dot diagrams.

Group 1 elements, hydrogen, lithium and sodium, each have a single dot, indicating one valence electron. Group 2 elements, beryllium and magnesium have two dots, for two valence electrons. Group 13 elements, boron and aluminum, have three dots and group 14 elements, carbon and silicon have four dots.

Now the group 15 elements, nitrogen and phosphorus, have five dots and because each side is now full, there is a pair of dots on one side, and single dots on the other three sides. Continuing with this pattern: Group 16 elements, oxygen and sulphur, have six dots, with two pairs and two single dots. Group 17 elements, fluorine and chlorine, with seven dots, has pairs on the left, right, and above, and a single dot below. Lastly, Group 18 elements, neon and argon, have eight dots, with pairs of dots on all four sides, representing a full valence shell.

Practice questions

Multiple-choice questions

1. Choose the correct Lewis structures of the molecule SO_3 .
 - a. The Lewis structure of sulphur trioxide has one sulphur atom and three oxygen atoms. Each oxygen is connected by a double bond to the sulphur atom, which is at the centre. The sulphur atom does not have a lone pair, while three oxygen atoms have two lone pairs.
 - b. The Lewis structure of sulphur trioxide has one sulphur atom and three surrounding oxygen atoms. Each oxygen is connected by a double bond to the sulphur atom. There are two lone pairs around the top oxygen and one pair around the other two oxygens, respectively.
 - c. The Lewis structure of sulphur trioxide has one sulphur atom and three surrounding oxygens. Each oxygen is connected by a double bond to the sulphur atom. Two lone pairs are around the bottom, and two oxygens and three electrons are around the top oxygen.
2. Choose the correct Lewis structures of the molecule COCl_2 .
 - a. The Lewis structure has carbon as the central atom with single bonds to chlorine atoms and double bonds to the oxygen atom. There are two lone pairs of electrons on the oxygen atom and each chlorine atom.
 - b. The Lewis structure has carbon as the central atom with single bonds to two chlorine atoms and one OH. There are three lone pairs of electrons on each chlorine atom. No electrons surround the OH.
 - c. The Lewis structure has carbon as the central atom with single bonds to chlorine atoms and double bonds to the oxygen atom. There are two lone pairs of electrons on the

- oxygen atom and three lone pairs on each chlorine atom.
3. Choose the correct Lewis structures of the molecule HF.
 - a. Hydrogen (H) and fluoride (F) are connected by a single bond. There are three lone pairs surrounding the fluoride.
 - b. Hydrogen (H) and fluoride (F) are connected by a double bond. There are two lone pairs surrounding the fluoride.
 - c. Hydrogen (H) and fluoride (F) are connected by a single bond. There are two lone pairs surrounding the fluoride.
 4. Choose the correct Lewis structures of the molecule SF₆.
 - a. The Lewis structure for sulphur hexafluoride involves a central sulphur atom bonded to six fluoride atoms with single bonds. There is one lone pair surrounding each fluoride atom.
 - b. The Lewis structure for sulphur hexafluoride involves a central sulphur atom bonded to six fluoride atoms with single bonds. There are three lone pairs surrounding each fluoride atom.
 - c. The Lewis structure for sulphur hexafluoride involves a central sulphur atom bonded to six fluoride atoms with single bonds. There are two lone pairs surrounding each fluoride atom.
 5. Choose the correct Lewis structures of the molecule CS₂.
 - a. The Lewis structure for carbon disulphide (CS₂) involves a central carbon double bonded to two sulphur atoms on each side. Each sulphur atom contains two lone pairs.
 - b. The Lewis structure for carbon disulphide (CS₂) involves a central carbon double bonded to two sulphur atoms on each side. Each sulphur atom contains three lone pairs.
 - c. The Lewis structure for carbon disulphide (CS₂) involves a central carbon double bonded to two sulphur atoms on each side. Each sulphur atom contains one lone pair.
 6. Choose the correct Lewis structures of the molecule SeF₄.
 - a. The Lewis structure of selenium tetrafluoride involves a selenium (Se) atom single-bonded to four fluoride (F)

- atoms. There are no extra electrons surrounding the fluoride or selenium.
- b. The Lewis structure of SeF_4 (selenium tetrafluoride) involves a selenium (Se) atom single-bonded to four fluoride (F) atoms. Each fluoride contains three lone pairs, while selenium contains one lone pair.
 - c. The Lewis structure of selenium tetrafluoride involves a selenium (Se) atom single-bonded to four fluoride (F) atoms. Each fluoride contains two lone pairs, while selenium contains no lone pairs.
7. Choose the correct Lewis structures of the molecule AlCl_3 .
- a. The Lewis structure for aluminium trichloride involves a central aluminium atom single-bonded to three chlorine atoms, with two lone pairs around each chloride. There are no extra electrons surrounding aluminium atoms.
 - b. The Lewis structure for aluminium trichloride involves a central aluminium atom single-bonded to three chlorine atoms, with one lone pair around each chloride. There are no extra electrons surrounding the aluminium atom.
 - c. The Lewis structure for aluminium trichloride involves a central aluminium atom single-bonded to three chlorine atoms, with three lone pairs around each chloride. There are no extra electrons surrounding aluminium atoms.
8. Choose the correct Lewis structures of the molecule Cl_2O .
- a. The Lewis structure for dichlorine monoxide involves a central oxygen atom with one lone pair and is double bonded to left chloride with one lone pair. Another chloride with three lone pairs is single-bonded to central oxygen.
 - b. The Lewis structure for dichlorine monoxide involves a central oxygen atom single-bonded to two chloride atoms, with the oxygen atom having two lone pairs and each chloride having two lone pairs.
 - c. The Lewis structure for dichlorine monoxide involves a

central oxygen atom single-bonded to two chloride atoms, with the oxygen atom having two lone pairs and each chloride having three lone pairs.

9. Choose the correct Lewis structures of the molecule N_2O .
- The Lewis structure for nitrous oxide involves two triple-bonded nitrogen atoms connected to one oxygen atom by a single bond. There are three lone pairs around oxygen. There is one lone pair around the left nitrogen.
 - The Lewis structure for nitrous oxide involves two triple-bonded nitrogen atoms connected to one oxygen atom by a double bond. There is one lone pair surrounding oxygen and the left nitrogen, respectively. The Lewis structure for nitrous oxide involves two triple-bonded nitrogen atoms connected to one oxygen atom by a double bond. There are two lone pairs around the oxygen.
10. Choose the correct Lewis structures of the molecule SiBr_4 .
- The Lewis structure of selenium tetrabromide involves a silicon (Si) atom double-bonded to four bromine (Br) atoms with three lone pairs around each bromine.
 - The Lewis structure of selenium tetrabromide involves a silicon (Si) atom single-bonded to four bromine (Br) atoms with three lone pairs around each bromine.
 - The Lewis structure of selenium tetrabromide involves a silicon (Si) atom single-bonded to four bromine (Br) atoms with two lone pairs around each bromine.

Solutions

Multiple-choice questions

1. (a)
2. (c)
3. (a)
4. (b)
5. (a)
6. (b)
7. (c)
8. (c)
9. (a)
10. (b)

12. 3.2 Ionic Bonds and Electronegativity

Figures

Figure 3.2.1: Electronegativity of Elements

A diagram of the s and p blocks on the periodic table displaying the electronegativity values of the elements. The diagram has two large arrows indicating increasing electronegativity across the periodic table—the first arrow begins at the top left and goes to the right; and the second arrow starts from the bottom left and goes vertical to the top of the table. The electronegativity values decrease from fluorine in group 7A which has the highest electronegativity at 3.98; moves down to francium in group 1 with the lowest at 0.7. Noble gases do not have an electronegativity value.

Figure 3.2.2: Formation of Magnesium Chloride

A lewis dot diagram of the the formation of magnesium chloride. To the left is a magnesium atom is shown with two valence electrons and two chlorine atoms each with seven valence electrons, indicated by lewis-dots. Red arrows show the movement of the two electrons from magnesium to the chlorine atoms. And on the right is the resulting ionic compound: the magnesium ion now labeled Mg^{2+} , indicating it has lost two electrons, and each chlorine atom is now a chloride ion, labeled Cl^- , indicating each has gained an extra electron, with a full outer shell depicted by lewis dots.

Practice questions

Multiple choice questions

1. Ca loses two electrons to acquire the electron configuration of Ar. What is the chemical symbol of the ion that it forms?
 - a. Ca^{2-}
 - b. Ca^{2+}
 - c. Ca
 - d. Ca^+
2. Which of the following statements is incorrect?
 - a. Metals tend to form cations.
 - b. Nonmetals tend to form anions.
 - c. The Octet rule states that atoms lose or gain electrons to acquire a completely filled valence shell that contains 8 electrons.
 - d. Noble gases are very reactive due to their electron configurations.
3. What is the electron configuration of Mg^{2+} ?
 - a. $1s^2 2s^2 2p^6$
 - b. $1s^2 2s^2 2p^6 3s^2$
 - c. $1s^2 2s^2 2p^6 3s^2 3p^6$
 - d. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
4. What is the electron configuration of F^- ?
 - a. $1s^2 2s^2 2p^5$
 - b. $1s^2 2s^2 2p^6$
 - c. $1s^2 2s^2 2p^6 3s^2 3p^6$
 - d. $1s^2$
5. What are the common ions formed by ^{29}Cu ?

- a. Cu^+ and Cu^{2+}
 - b. Cu^+ and Cu^{3+}
 - c. Cu^{2+} and Cu^{3+}
 - d. Cu^+ , Cu^{2+} and Cu^{3+}
6. What is the symbol of the ion formed by Beryllium losing two electrons from the valence shell?
- a. Be^+
 - b. Be^{2+}
 - c. Ba^{2+}
 - d. Ba^+
7. Which of the following options incorrectly names the given polyatomic ion?
- a. CO_3^{2-} = Carbonate ion
 - b. CH_3COO^- = Acetate ion
 - c. CN^- = Cyanide ion
 - d. MnO_4^- = Manganese ion
8. What is the formula of the compound formed by NH_4^+ and Cl^- ions?
- a. NH_4Cl
 - b. NHCl_4
 - c. ClNH_4
 - d. NHCl
9. What is the chemical formula for calcium carbonate?
- a. CaCO_3
 - b. CaCO
 - c. CaO
 - d. CaHCO_3
10. Which of the following options matches the name and the chemical formula?
- a. Ammonium carbonate: $(\text{NH}_4)_2\text{CO}_3$
 - b. Magnesium hydroxide: MgOH
 - c. Silver nitrate: AgNO
 - d. Sodium hypochlorite: NaClO
11. Which of the following options gives the correct name for

AuCl_3 ?

- a. Gold chloride
- b. Gold chloride (III)
- c. Gold (III) chloride
- d. Gold (III) chloride (III)

Solutions

- 1. b
- 2. d
- 3. a
- 4. b
- 5. a
- 6. b
- 7. d
- 8. a
- 9. a
- 10. a
- 11. c

13. 3.3 Covalent Bonds and Polarity

Figures

Figure 3.3.2: Electronegativity of Elements

A diagram of the s and p blocks on the periodic table displaying the electronegativity values of the elements. The diagram has two large arrows indicating increasing electronegativity across the periodic table—the first arrow begins at the top left and goes to the right; and the second arrow starts from the bottom left and goes vertical to the top of the table. The electronegativity values decrease from fluorine in group 7A which has the highest electronegativity at 3.98; moves down to francium in group 1 with the lowest at 0.7. Noble gases do not have an electronegativity value.

Practice questions

Multiple-choice questions

1. Determine the molecular formulas of the H_xS by completing subscripts $x=?$
 - a. 2
 - b. 4
 - c. 6

- d. 8
- 2. Determine the molecular formulas of the CCl_z by completing subscripts $z=?$
 - a. 4
 - b. 2
 - c. 1
 - d. 6
- 3. Determine the molecular formulas of the PF_w by completing subscripts $w=?$
 - a. 1
 - b. 2
 - c. 3
 - d. 4
- 4. Determine whether the following bonds are ionic, polar covalent, or covalent: between S and Cl.
 - a. ionic
 - b. polar covalent
 - c. covalent
- 5. Determine whether the following bonds are ionic, polar covalent, or covalent: between N and H.
 - a. ionic
 - b. polar covalent
 - c. covalent
- 6. Determine whether the following bonds are ionic, polar covalent, or covalent: between C and Br.
 - a. ionic
 - b. polar covalent
 - c. covalent
- 7. Determine whether the following bonds are ionic, polar covalent, or covalent: between Li and F.
 - a. ionic
 - b. polar covalent
 - c. covalent

8. Carbon, oxygen, nitrogen, halides, and hydrogen are frequently found in molecular compounds. They generally maintain a consistent bonding pattern to acquire an octet in the valence shell. Which of the following statements is not correct regarding the general bonding pattern of the atoms mentioned above?
- Carbon forms four covalent bonds
 - Nitrogen forms three covalent bonds and has one lone pair of electrons
 - Oxygen forms two covalent bonds and has two lone pairs of electrons
 - Hydrogen forms one covalent bond and has one lone pair of electrons
9. Which of the following pairs of compounds only contain covalent bonds?
- CO, NaCl
 - NCl₃, CH₄
 - CH₄, CaCl₂
 - NaF, FeCl₃
10. The covalent bond that forms between boron and nitrogen is polar covalent. Which of the following options correctly shows the dipoles in the bond?
- B δ^+ – N δ^-
 - B δ^- – N δ^+
 - B δ^+ – N δ^+
 - B δ^- – N δ^-
11. Choose the correct Lewis structure for the Cl₂ compound.
- The chlorine gas Lewis structure consists of two chlorine atoms linked by a triple bond with two lone pairs on each chlorine.
 - The chlorine gas Lewis structure consists of two chlorine atoms linked by a single bond with three lone pairs on each chlorine.
 - The chlorine gas Lewis structure consists of two chlorine

atoms linked by a single bond with two lone pairs on each chlorine.

12. Choose the correct Lewis structure for the HI compound.
 - a. The Lewis structure for hydrogen iodide involves a hydrogen atom single-bonded to an iodine atom with three lone pairs around the iodine atom.
 - b. The Lewis structure for hydrogen iodide involves a hydrogen atom single-bonded to an iodine atom with no extra electrons around both iodine and hydrogen atoms.
 - c. The resulting Lewis structure for hydrogen iodide involves a hydrogen atom single-bonded to an iodine atom with three lone pairs around both hydrogen and iodine atoms.
13. Choose the correct Lewis structure for the PCl_3 compound.
 - a. The Lewis structure for phosphorus trichloride involves a central phosphorus atom single-bonded to three chlorine atoms with three lone pairs around each chlorine atom.
 - b. The Lewis structure for phosphorus trichloride involves a central phosphorus atom single-bonded to three chlorine atoms with three lone pairs around each chlorine atom and one lone pair around the central phosphorus.
 - c. The Lewis structure for phosphorus trichloride involves a central phosphorus atom single-bonded to two chlorine atoms with one lone pair around each chlorine atom and triple-bonded to the third chlorine with one lone pair around it.

Short-answer questions

1. Determine the molecular formulas of HBr_v by completing subscripts v =?
2. Name the following molecular compounds:
 - a. BBr_3

- b. CS_2
- c. S_2Cl_2
- d. N_2O_3

Solutions

Multiple-choice questions

- 1. a
- 2. a
- 3. c
- 4. b
- 5. b
- 6. c
- 7. a
- 8. d
- 9. b
- 10. a
- 11. b
- 12. a
- 13. b

Short-answer questions

- 1. one/ 1
- 2.
 - a. Boron tribromide
 - b. Carbon disulfide
 - c. Disulphur dichloride

d. Dinitrogen trioxide

14. 3.4 Metallic Bonds

Practice questions

Multiple-choice questions

1. Which of the following is a property of metals due to metallic bonding?
 - a. They are malleable and ductile.
 - b. They have low melting points.
 - c. They have low electrical conductivity.
2. In metallic bonding, what is the role of the “sea of electrons”?
 - a. It generates heat.
 - b. It provides electrical conductivity.
 - c. It forms covalent bonds.
 - d. It generates electricity.
3. At room temperature, most metals are either a:
 - a. Solid
 - b. Gas
 - c. Liquid
4. What does malleable mean?
 - a. Look shiny
 - b. Join easily
 - c. Break easily
 - d. Able to be shaped
5. Why do metals conduct?
 - a. The electrons are held tightly within the lattice.
 - b. The electrons are delocalised and able to move.
 - c. The electrons are shared between two metal ions.

- d. The electrons show negative charges, which are easier to conduct heat.

True/false question

6. Metallic bonding is the bond between two metals.

Solutions

Multiple-choice questions

1. a
2. b
3. a
4. d
5. b

True/false questions

6. True

15. 3.5 Intermolecular Forces

Practice questions

Multiple-choice questions (There is/ are one or more correct answer(s) for each question)

1. Butane will experience what type(s) of intermolecular forces?
 - a. Dispersion Forces
 - b. Hydrogen bonding Forces
 - c. Dipole-dipole interactions Forces
 - d. London Forces
2. Hydrochloric Acid HCl will experience what type(s) of intermolecular forces?
 - a. Hydrogen Bonding Force
 - b. Dipole-Dipole Interactions Force
 - c. (London) Dispersion Forces
3. Which of the following options is the odd one out? Water, Ammonia, Pentane or Hydrofluoric acid?
 - a. Water
 - b. Ammonia
 - c. Pentane
 - d. Hydrofluoric acid

True/false question

4. Molecules that are non-polar and do not have permanent dipoles are unable to generate intermolecular forces.

Solutions

Multiple-choice questions

1. a and d. London dispersion forces are also known as dispersion forces and London forces.
2. a; b and c
3. c

True/false question

4. False

16. 3.6 Predicting Molecular Shape

Figures

Figure 3.6.7: Methane ammonia and water

Three molecular diagrams of methane, ammonia and water, each with different shapes due to varying numbers of lone pairs.

Methane (CH_4) has no lone pairs, and is a tetrahedral structure drawn as a central carbon surrounded by four hydrogen. Single bonds are in the frontal plane and the back plane, represented by a thick wedge and dashes, respectively.

In the center, ammonia (NH_3) has a trigonal pyramidal shape with one lone pair above its chemical symbol. Single bonds are again in the frontal plane and the back plane, represented by a thick wedge and dashes, respectively.

On the right, water (H_2O) has a bent structure due to two lone pairs on the oxygen atom, and a single wedge for one of the hydrogen bonds with other being drawn as a single line.

Figure 3.6.8: Table of Molecular Geometry

Table illustrating VSEPR theory molecular geometries based on the number of electron groups and lone pairs around a central atom. Each cell contains a 3D model, and each model uses a dark sphere to represent the central atom, red spheres for bonded atoms, and light spheres with dots for lone pairs. The geometries are shown

transitioning from fully bonded atoms to those with increasing lone pairs, altering the molecule's shape.

In the table, the heading for the first column is Number of Electron Groups.

- In row one, with 2 electron groups: a linear molecular shape with 0 lone pairs.
- In row two with 3 electron groups: a trigonal planar with 0 lone pairs, then in column 2 a either angular or bent in shape, with 1 lone pair.
- In row three with 4 electron groups: a tetrahedral with 0 lone pairs, then in column two, a trigonal pyramidal with 1 lone pair, then in column 3, an angular or bent shape with 2 lone pairs.
- In row four with 5 electron groups: a trigonal bipyramidal with 0 lone pairs, then in column 2, a seesaw with 1 lone pair, then in column 3, a t-shape with 2 lone pairs, then in column 4, linear with 3 lone pairs.
- In row five with 6 electron groups: an octahedral with 0 lone pairs, then in column 2, square pyramidal with 1 lone pair, then in column 3, a square planar with 2 lone pairs, then in column 4, a t-shape with 3 lone pairs, and finally in column 5, a linear shape with 4 lone pairs.

Figure 3.6.9

A Lewis dot diagram of carbon dioxide, with a central carbon atom double-bonded to two oxygen atoms, each with two pairs of lone electrons. Arrows indicate the polar nature of the carbon-oxygen bonds with the arrow heads pointing towards oxygen, denoting the negative pole, and tails at the carbon, the positive pole. Large cross marks over the arrows illustrate that the opposing polarities cancel out leading to no net dipole moment for the molecule, rendering it nonpolar.

Figure 3.6.10

Four lewis dot diagrams depicting the polarity of water molecules (H_2O) from top to bottom:

1. An oxygen atom in the center with two lone pairs of electrons and single bonds to two hydrogen atoms. Two crossed arrows above the oxygen atom point towards the hydrogen atoms, indicating the polar nature of the O-H bonds.
2. An oxygen atom at the center bonded to two hydrogen atoms. The diagram includes four arrows, two either side of the oxygen molecule, pointing inwards from the hydrogen atoms and again towards the oxygen, representing the electronegativity difference.
3. An oxygen atom at the center bonded to two hydrogen atoms. The diagram shows the molecule's angular shape with two lone pairs of electrons on the oxygen. Horizontal arrows with plus and minus signs indicate the horizontal components of the polar covalent bonds, which are canceled out due to their opposite directions. The vertical components, represented by vertical arrows, do not cancel, resulting in a net dipole moment pointing upwards.
4. An oxygen atom at the center bonded to two hydrogen atoms. A large vertical arrow above the oxygen indicates the net dipole moment, resulting from the molecule's bent shape which doesn't cancel out the vertical components of the polar bonds, thus making water polar.

Practice questions

Multiple-choice questions

1. Predict the geometry of Phosphorous trifluoride PF_3 . Is this molecule polar or non-polar? (Hint: Draw the Lewis Structure to determine the number of electron-dense areas.)
 - a. Phosphorous trifluoride is a compound with 4 electronegative regions and 1 lone pair, making it a trigonal pyramid. It is a polar compound.
 - b. Phosphorous trifluoride is a compound with 3 electronegative regions and 2 lone pairs, making it a trigonal pyramid. It is a polar compound.
 - c. Phosphorous trifluoride is a compound with 4 electronegative regions and 2 lone pairs, making it a trigonal pyramid. It is a non-polar compound.
 - d. Phosphorous trifluoride is a compound with 4 electronegative regions and 1 lone pair, making it a trigonal pyramid. It is a non-polar compound.

Short-answer questions

1. Fill in the missing words/numbers.
 SCl_2 Sulphur dichloride is a compound with _____ electronegative regions and _____ lone pairs, making it angular or bent. It is a _____ compound.
2. Fill in the missing numbers in the first two boxes. Write down the polarity in the last two boxes.
 SiBr_4 Silicon tetrabromide is a compound with _____ electronegative regions and _____ lone pairs, making it a _____ compound. It is _____.

Solutions

Multiple-choice questions

1. a

Short-answer questions

Fill in the missing words/numbers

1. 4/ four; 2/ two; polar
2. 4/ four; 0/ zero; tetrahedral/ non-polar; non-polar/
tetrahedral

17. 4.1 Chemical Equations

Practice questions

1. From the statement 'nitrogen and hydrogen react to produce ammonia,' identify the reactants and products.
 - a. Reactants: ammonia and nitrogen. Products: hydrogen.
 - b. Reactants: nitrogen. Products: ammonia and hydrogen.
 - c. Reactants: ammonia. Products: nitrogen and hydrogen.
 - d. Reactants: nitrogen and hydrogen. Products: ammonia.
2. From the statement 'sodium metal reacts with water to produce sodium hydroxide and hydrogen,' identify the reactants and the products.
 - a. Reactants: sodium and water. Products: sodium hydroxide and hydrogen.
 - b. Reactants: sodium, hydrogen and water. Products: sodium hydroxide.
 - c. Reactants: sodium hydroxide and hydrogen. Products: sodium and water.
 - d. Reactants: sodium hydroxide. Products: sodium, hydrogen and water.
3. What does the arrow in a chemical equation represent?
 - a. The conversion of reactants into products.
 - b. The reactants and products are unrelated.
 - c. The addition of reactants.

- d. An equal sign.
- 4. In a balanced chemical equation, which of the following is conserved?
 - a. Mass and volume.
 - b. Mass only.
 - c. Volume only.
 - d. Mass and moles.
- 5. Which of the following statements is true about a chemical equation?
 - a. The physical state of the reactants and products is written as a superscript after each substance.
 - b. The coefficients represent the ratio of amounts of reactants and products.
 - c. The reactants and products are always in the same physical state.
 - d. The conditions required for the reaction to occur can be written before the reactants.
- 6. When balancing a chemical equation, which of the following is adjusted to ensure the conservation of mass?
 - a. Subscripts.
 - b. Nothing needs to be adjusted.
 - c. The state of matter.
 - d. Coefficients.
- 7. What does a chemical equation represent?
 - a. A summary of the periodic table.
 - b. A list of chemical elements.
 - c. A recipe for a chemical reaction.
 - d. A list of chemical compounds.
- 8. Which of the following is a reactant in the chemical equation for photosynthesis:

$$6 \text{CO}_2 + 6 \text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2$$
 - a. CO_2

- b. O_2
 - c. $\text{C}_6\text{H}_{12}\text{O}_6$
 - d. H_2
9. From the statement 'magnesium hydroxide reacts with nitric acid to produce magnesium nitrate and water,' identify the reactants and the products.
- a. Reactants: magnesium nitrate and water.
Products: magnesium hydroxide and nitric acid.
 - b. Reactants: magnesium hydroxide. Products: magnesium nitrate and water.
 - c. Reactants: magnesium hydroxide and water.
Products: magnesium nitrate and nitric acid.
 - d. Reactants: magnesium hydroxide and nitric acid.
Products: magnesium nitrate and water.
10. From the statement 'propane reacts with oxygen to produce carbon dioxide and water,' identify the reactants.
- a. Oxygen only.
 - b. Carbon dioxide and water.
 - c. Propane only.
 - d. Propane and oxygen.

Answers

- 1. d
- 2. a
- 3. a
- 4. b
- 5. b
- 6. d
- 7. c
- 8. a

9. d

10. d

18. 4.2 Writing and Balancing Chemical Equations

Practice questions

1. What is the balanced chemical equation for the reaction between hydrogen and water to form HCl ?
 - a. $\text{H}_2 + \text{Cl}_2 \rightarrow 2 \text{HCl}$
 - b. $2 \text{H}_2 + \text{Cl}_2 \rightarrow 2 \text{HCl}$
 - c. $\text{H}_2 + \text{Cl}_2 \rightarrow 4 \text{HCl}$
 - d. $2 \text{H}_2 + 2 \text{Cl}_2 \rightarrow 2 \text{HCl}$
2. Ethane, C_2H_6 reacts with oxygen to make carbon dioxide and water. What is the balanced chemical equation for this reaction?
 - a. $4 \text{C}_2\text{H}_6 + 7 \text{O}_2 \rightarrow 4 \text{CO}_2 + 6 \text{H}_2\text{O}$
 - b. $2 \text{C}_2\text{H}_6 + 7 \text{O}_2 \rightarrow 4 \text{CO}_2 + 12 \text{H}_2\text{O}$
 - c. $2 \text{C}_2\text{H}_6 + 5 \text{O}_2 \rightarrow 4 \text{CO}_2 + 6 \text{H}_2\text{O}$
 - d. $2 \text{C}_2\text{H}_6 + 7 \text{O}_2 \rightarrow 2 \text{CO}_2 + 6 \text{H}_2\text{O}$
 - e. $2 \text{C}_2\text{H}_6 + 7 \text{O}_2 \rightarrow 4 \text{CO}_2 + 6 \text{H}_2\text{O}$
3. Fill in the missing coefficient(s) to balance the equation:
..... $\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$
4. Balance the following equation by filling in the blanks:
..... $\text{NaClO}_3 \rightarrow \text{NaCl} + 3 \text{O}_2$
5. Balance the following equation by filling in the blanks:
..... $\text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3$
6. Balance the following equation by filling in the blanks:
 $\text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$

7. Balance the following equation by filling in the blanks:
 $\text{Ca}(\text{OH})_2 + \dots\dots\dots \text{HCl} \rightarrow \text{CaCl}_2 + \dots\dots\dots \text{H}_2\text{O}$
8. The balanced equation for the reaction between silver nitrate and magnesium chloride is:
 $2 \text{AgNO}_3 + \text{MgCl}_2 \rightarrow 2 \text{AgCl} + \text{Mg}(\text{NO}_3)_2$
- Is the above statement true or false?
9. What is the balanced chemical equation for the reaction between hydrochloric acid and calcium carbonate to form carbon dioxide, calcium chloride and water?
- $2 \text{HCl} + 2 \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$
 - $4 \text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{CO}_2 + 2 \text{H}_2\text{O}$
 - $\text{HCl} + 2 \text{CaCO}_3 \rightarrow 2 \text{CaCl}_2 + 2 \text{CO}_2 + \text{H}_2\text{O}$
 - $2 \text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$
10. The balanced chemical equation for the reaction between ammonia gas and chlorine gas to form hydrazine is:
 $4 \text{NH}_3 + \text{Cl}_2 \rightarrow 2 \text{N}_2\text{H}_4 + 2 \text{NH}_4\text{Cl}$. Is this statement true or false?

Answers

- a
- e
- 2
- 2, 2
- 4, 3, 2
- 2, 2
- 2, 2
- True
- d
- True

19. 4.3 Classes of Chemical Reactions: Single and Double Displacement Reactions

Practice questions

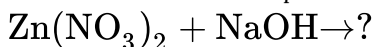
1. What are the general characteristics that help you recognise single-replacement reactions?
 - a. Only cations are swapped.
 - b. One element replaces another element in a compound.
 - c. Only anions are swapped.
 - d. Parts of two ionic compounds are exchanged, making two new compounds.
2. What are the general characteristics that help you recognise double-replacement reactions?
 - a. The exchange of the cations or the anions of two ionic compounds, making two new compounds.
 - b. They can occur between any given reactants.
 - c. The generation of a new element and a compound.
 - d. One element replaces another element in a compound.
3. Assuming that the following single-replacement reaction occurs, predict the products of the reaction: $\text{Zn} + \text{Fe}(\text{NO}_3)_2 \rightarrow ?$
 - a. $\text{Zn} + \text{Fe} + \text{NO}_2$

- b. $\text{FeZn} + \text{N}_2$
 - c. $\text{Zn} + \text{Fe}(\text{NO}_3)$
 - d. $\text{Zn}(\text{NO}_3)_2 + \text{Fe}$
4. Assuming that the following single-replacement reaction occurs, predict the products of the reaction: $\text{Li} + \text{MgSO}_4 \rightarrow ?$
 - a. LiSO_4 and MgSO_4
 - b. Li and Mg
 - c. Li_2SO_4 and Mg
 - d. MgSO_4 and Li
 5. Assuming that the following single-replacement reaction occurs, predict the products and write the balanced chemical equation:
 $\text{F}_2 + \text{FeI}_3 \rightarrow ?$
 - a. $3 \text{F}_2 + 2 \text{FeI}_3 \rightarrow 3 \text{F}_2 + 2 \text{FeF}_3$
 - b. $3 \text{F}_2 + 2 \text{FeI}_3 \rightarrow 3 \text{I}_2 + 2 \text{FeF}_3$
 - c. $\text{F}_2 + 2 \text{FeI}_3 \rightarrow 3 \text{I}_2 + 2 \text{FeF}_3$
 - d. $3 \text{F}_2 + 2 \text{FeI}_3 \rightarrow 3 \text{I}_2 + 2 \text{Fe} + \text{F}_2$
 6. Assuming that the following single-replacement reaction occurs, predict the products and write a balanced chemical equation: $\text{Al} + \text{NiBr}_2 \rightarrow ?$
 7. Decide whether the following statement is correct regarding the single-replacement reaction between Sn and H_2SO_4 . Sn reacts with H_2SO_4 to produce SnSO_4 and H_2 . The balanced chemical equation for this reaction is $\text{Sn} + \text{H}_2\text{SO}_4 \rightarrow \text{SnSO}_4 + \text{H}_2$.
 8. Use the periodic table or the activity series to predict if the following single-replacement reaction will occur and, if so, what is the balanced chemical equation? $\text{FeCl}_2 + \text{Br}_2 \rightarrow ?$

9. Use the periodic table or the activity series to predict if the following single-replacement reaction will occur and, if so, what is the balanced chemical equation? $\text{Fe}(\text{NO}_3)_3 + \text{Al} \rightarrow ?$

- a. $3 \text{Fe}(\text{NO}_3)_3 + 2 \text{Al} \rightarrow 2 \text{Al}(\text{NO}_3)_3 + 3 \text{Fe}$
- b. $\text{Fe}(\text{NO}_3)_3 + \text{Al} \rightarrow \text{Al}(\text{NO}_3)_3 + \text{Fe}$
- c. $\text{Fe}(\text{NO}_3)_3 + \text{Al} \rightarrow \text{Al} + \text{Fe}$
- d. No reaction occurs

10. Assuming that the following double-replacement reaction occurs, predict the products and write the balanced chemical equation:



- a. $\text{Zn}(\text{NO}_3)_2 + \text{NaOH} \rightarrow \text{Zn}(\text{OH}_2) + \text{NaNO}_3$
- b. $\text{Zn}(\text{NO}_3)_2 + 2 \text{NaOH} \rightarrow \text{Zn}(\text{OH}_2) + 2 \text{NaNO}_3$
- c. $\text{Zn}(\text{NO}_3)_2 + 2 \text{NaOH} \rightarrow \text{Zn}(\text{OH}_2) + 2 \text{Na}$
- d. $\text{Zn}(\text{NO}_3)_2 + 2 \text{NaOH} \rightarrow \text{Zn} + 2 \text{NaNO}_3$

11. Assuming that the following double-replacement reaction occurs, predict the products and write each balanced chemical equation:



- a. $2 \text{HCl} + \text{Na}_2\text{S} \rightarrow 2 \text{NaCl} + \text{H}_2\text{S}$
- b. $2 \text{HCl} + \text{Na}_2\text{S} \rightarrow 2 \text{NaCl} + \text{H}_2$
- c. $2 \text{HCl} + \text{Na}_2\text{S} \rightarrow 2 \text{NaH} + \text{Cl}_2$
- d. $\text{HCl} + \text{Na}_2\text{S} \rightarrow 2 \text{NaCl} + \text{H}_2\text{S}$

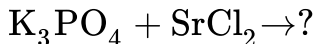
12. Assuming that each double-replacement reaction occurs, predict the products and write the balanced chemical equation:



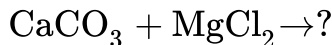
13. Assuming that the following double-replacement reaction occurs, predict the products and write the balanced chemical equation:



14. Use the solubility rules to predict if the following double-replacement reaction will occur and, if so, what is the balanced chemical equation?



- $2 \text{K}_3\text{PO}_4 + 3 \text{SrCl}_2 \rightarrow \text{Sr}_3(\text{PO}_4)_2(\text{s}) + 6 \text{KCl}$
 - $\text{K}_3\text{PO}_4 + \text{SrCl}_2 \rightarrow \text{Sr}_3(\text{PO}_4)_2(\text{s}) + \text{KCl}$
 - No reaction occurs.
 - $2 \text{K}_3\text{PO}_4 + 3 \text{SrCl}_2 \rightarrow \text{Sr}(\text{s}) + 6 \text{KCl}$
15. Use the solubility rules to predict if the following double-replacement reaction will occur and, if so, what is the balanced chemical equation?



- $\text{CaCO}_3 + \text{MgCl}_2 \rightarrow \text{CaCl}_2 + \text{Mg}$
 - $\text{CaCO}_3 + 2 \text{MgCl}_2 \rightarrow \text{CaCl}_2 + 2 \text{MgCO}_3$
 - $\text{CaCO}_3 + \text{MgCl}_2 \rightarrow \text{CaCl}_2 + \text{MgCO}_3$
 - No reaction occurs
16. Decide whether the following statement is true. Na_2CO_3 is soluble in water. *Hint: use solubility rules to predict.*
17. Decide whether the following statement is true. CaCl_2 is soluble in water. *Hint: use solubility rules to predict.*
18. A precipitation reaction will occur when an aqueous solution of AgNO_3 reacts with an aqueous solution of Na_2CO_3 .

Answers

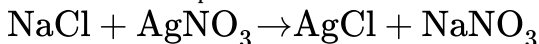
- b
- a

3. d
4. c
5. b
6. $2 \text{Al} + 3 \text{NiBr}_2 \rightarrow 2 \text{AlBr}_3 + 3 \text{Ni}$
7. True
8. b
9. c
10. b
11. a
12. $\text{Pb}(\text{NO}_3)_2 + 2 \text{KBr} \rightarrow \text{PbBr}_2 + 2 \text{KNO}_2$
13. $\text{K}_2\text{O} + \text{MgCO}_3 \rightarrow \text{K}_2\text{CO}_3 + \text{MgO}$
14. a
15. d
16. True
17. True
18. True

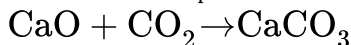
20. 4.4 Composition, Decomposition, and Combustion Reactions

Practice questions

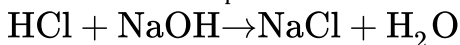
1. Is the following statement true or false? The following reaction is a composition reaction:



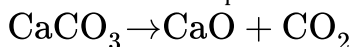
2. Is the following statement true or false? The following reaction is a composition reaction:



3. Is the following statement true or false? The following reaction is a decomposition reaction:



4. Is the following statement true or false? The following reaction is a decomposition reaction:



5. Which of the following is a combustion reaction?

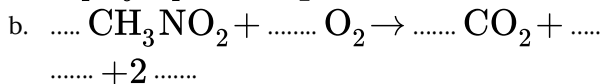
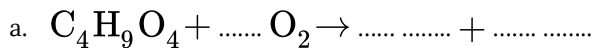
- a. $6 \text{C} + 3 \text{H}_2 \rightarrow \text{C}_6\text{H}_6$
- b. $\text{H}_2\text{SO}_3 \rightarrow \text{H}_2\text{O} + \text{SO}_2$
- c. $\text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{CO}_2 \rightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O}$
- d. $\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3$

6. Which of the following is a combustion reaction?

- a. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
- b. $2 \text{Fe}_2\text{S}_3 + 9 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3 + 6 \text{SO}_2$
- c. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$



7. Complete and balance the following combustion reactions.



Answers

1. False

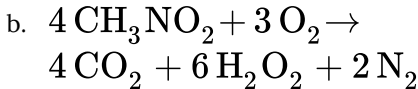
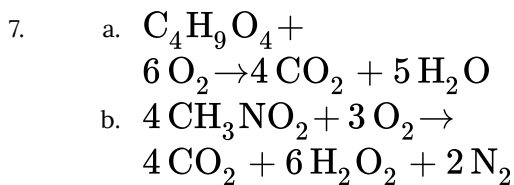
2. True

3. False

4. True

5. c

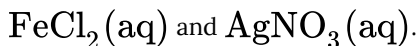
6. b



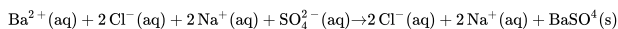
2I. 4.5 Ionic Equations

Practice questions

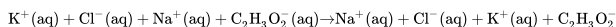
1. What is the chemical equation that represents NaBr dissociating in water?
 - a. $\text{NaBr}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{Br}^-(\text{aq})$
 - b. $\text{NaBr}(\text{s}) \rightarrow \text{Na}_2^+(\text{aq}) + \text{Br}_2^-(\text{aq})$
 - c. $\text{NaBr}(\text{s}) \rightarrow \text{NaOH}(\text{aq}) + \text{BrH}(\text{aq})$
 - d. $\text{NaBr}(\text{s}) \rightarrow \text{NaBr}(\text{aq})$
2. What is the chemical equation that represents $\text{SrCl}_2(\text{s})$ dissociation in water?
 - a. $\text{SrCl}_2(\text{s}) \rightarrow \text{Sr}_2^{2+}(\text{aq}) + 2 \text{Cl}_2^-$
 - b. $\text{SrCl}_2(\text{s}) \rightarrow \text{Sr}^{2+}(\text{aq}) + 2 \text{Cl}^-$
 - c. $\text{SrCl}_2(\text{s}) \rightarrow \text{SrCl}_2(\text{aq})$
 - d. $\text{SrCl}_2(\text{s}) \rightarrow \text{Sr}^{2+}(\text{aq}) + \text{Cl}^-$
3. Which of the following represents the chemical equation of $(\text{NH}_4)_3\text{PO}_4(\text{s})$ dissociating in water?
 - a. $(\text{NH}_4)_3\text{PO}_4(\text{s}) \rightarrow (\text{NH}_4)_3\text{PO}_4(\text{aq}) + \text{H}_2\text{O}$
 - b. $(\text{NH}_4)_3\text{PO}_4(\text{s}) \rightarrow 3 \text{NH}_4^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
 - c. $(\text{NH}_4)_3\text{PO}_4(\text{s}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
 - d. $(\text{NH}_4)_3\text{PO}_4(\text{s}) \rightarrow (\text{NH}_4)_3\text{PO}_4(\text{aq})$
4. Write the chemical equation that represents $\text{Fe}(\text{C}_2\text{H}_3\text{O}_2)_3(\text{s})$ dissociating in water.
5. Write the complete ionic equation for the reaction of



6. Is the following statement true? This is the complete ionic equation for the reaction of $\text{BaCl}_2(\text{aq})$ and $\text{Na}_2\text{SO}_4(\text{aq})$:

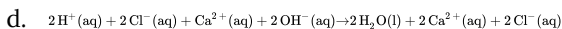
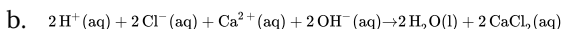


7. Is the following statement true? This is the complete ionic equation for the reaction of $\text{KCl}(\text{aq})$ and $\text{NaC}_2\text{H}_3\text{O}_2(\text{aq})$:

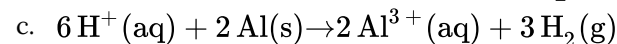
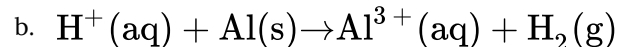
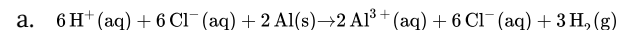
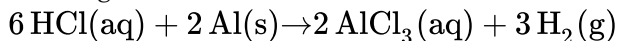


8. Write the net ionic equation for the reaction of $\text{FeCl}_2(\text{aq})$ and $\text{AgNO}_3(\text{aq})$.

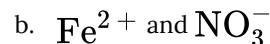
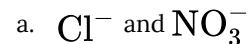
9. What is the balanced net ionic equation for the reaction between HCl and $\text{Ca}(\text{OH})_2(\text{aq})$?



10. What is the balanced net ionic equation for the following reaction:



11. Identify the spectator ions in the reaction between



- c. Ag^+ and NO_3^-
 - d. Ag^+ and Cl^-
12. Identify the spectator ions in the following reaction:
 $\text{BaCl}_2(\text{aq})$ and $\text{Na}_2\text{SO}_4(\text{aq})$.
- a. Na^+ and Cl^-
 - b. Ba^{2+} and Cl^-
 - c. Ba^{2+} and SO_4^{2-}
 - d. Na^+ and SO_4^{2-}

Answers

1. a
2. b
3. b
4. $\text{Fe}(\text{C}_2\text{H}_3\text{O}_2)_3(\text{s}) \rightarrow \text{Fe}^{3+}(\text{aq}) + 3 \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$
5. $\text{Fe}^{2+}(\text{aq}) + 2 \text{Cl}^-(\text{aq}) + 2 \text{Ag}^+(\text{aq}) + 2 \text{NO}_3^-(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2 \text{NO}_3^-(\text{aq}) + 2 \text{AgCl}(\text{s})$
6. True
7. True
8. $2 \text{Cl}^-(\text{aq}) + 2 \text{Ag}^+(\text{aq}) \rightarrow 2 \text{AgCl}(\text{s})$
9. a
10. c
11. b
12. a

22. 4.6 Oxidation-Reduction Reactions

Practice questions

1. Is this statement true? The following reaction is a redox reaction: $2 \text{K(s)} + \text{Br}_2 \text{(l)} \rightarrow 2 \text{KBr(s)}$
2. In the reaction $\text{Ca(s)} + \text{O}_2 \text{(g)} \rightarrow 2 \text{CaO}$, indicate which element/compound has lost electrons and what element/compound has gained electrons.
3. In the reaction $2 \text{Li(s)} + \text{O}_2 \text{(g)} \rightarrow \text{Li}_2 \text{O}_2 \text{(s)}$, indicate which element/compound has been oxidised and which element/compound has been reduced.
4. What is oxidation? *Select all applicable answers.*
 - a. The loss of electrons
 - b. The gaining of electrons
 - c. An increase in the oxidation number
 - d. A decrease in the oxidation number
5. What is reduction? *Select all applicable answers.*
 - a. Increase in oxidation number
 - b. Decrease in the oxidation number
 - c. Gain of electrons
 - d. Loss of electrons
6. Assign oxidation numbers to each atom in each substance:
 - a. P in P_4
 - b. S and O in SO_2
 - c. S and O in SO_2^{2-}
 - d. Ca, N and O in $\text{Ca}(\text{NO}_3)_2$

7. Assign oxidation numbers to each atom in each substance:
- C and O in CO
 - C and O in CO₂
 - Ni and Cl in NiCl₂
 - Ni and Cl in NiCl₃
8. Identify what is being oxidised and reduced in the redox equation $2\text{NO} + \text{Cl}_2 \rightarrow 2\text{NOCl}$ by assigning oxidation numbers to the atoms. N is being, and Cl is being
9. Identify what is being oxidised and reduced in the redox equation $2\text{KrF}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{Kr} + 4\text{HF} + \text{O}_2$ by assigning oxidation numbers to the atoms. O is being, and Kr is being
10. Identify the oxidising and reducing agents in the following reactions:
- $\text{Cu(s)} + \text{Pt}^{2+}(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{Pt(s)}$
 - $2\text{Mg(s)} + \text{CO}_2(\text{g}) \rightarrow 2\text{MgO(s)} + \text{C(s)}$
11. Which of the following is the correct balanced net ionic equation for the reaction shown below?
- $$\text{Cr}_2\text{O}_7^{2-} + \text{Fe}^{2+} \rightarrow \text{Cr}^{3+} + \text{Fe}^{3+}$$
- $14\text{H}^+ + \text{Cr}_2\text{O}_7^{2-} + \text{Fe}^{2+} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + \text{Fe}^{3+}$
 - $14\text{H}^+ + \text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} \rightarrow 2\text{Cr}^{3+} + 14\text{H}_2\text{O} + 6\text{Fe}^{3+}$
 - $14\text{H}^+ + \text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 6\text{Fe}^{3+}$
 - $14\text{H}^+ + \text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 6\text{Fe}^{3+}$
12. Which of the following is the correct balanced net ionic equation for the reaction shown below?
- $$\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2\text{KI}(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbI}_2(\text{s})$$
- $\text{Pb}^{2+}(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$
 - $\text{Pb}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{K}^+ + 2\text{I}^-(\text{aq}) \rightarrow 2\text{K}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{PbI}_2(\text{s})$

- c. $\text{Pb}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$
 d. $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2\text{KI}(\text{aq}) \rightarrow 2\text{KNO}_3(\text{aq}) + \text{PbI}_2$

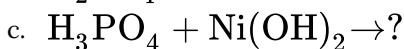
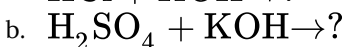
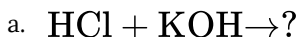
Answers

- True
- Lost electrons - **Ca** and gained electrons - **O₂**
- Oxidised - **Li**, Reduced - **O₂**
- a and c
- b and c
- Oxidation numbers are as follows:
 - P** = 0
 - S** = +4 and **O** = -2
 - S** = +2 and **O** = -2
 - Ca** = +2, **N** = +5, and **O** = -2
- Oxidation numbers are as follows:
 - C** = +2 and **O** = -2
 - C** = +4 and **O** = -2
 - Ni** = +2 and **Cl** = -1
 - Ni** = +3 and **Cl** = -1
- Oxidised, reduced
- Oxidised, reduced
- Reducing and oxidising agents are as follows:
 - Reducing agent **Cu** and oxidising agent **Pt²⁺**
 - Reducing agent **Mg** and oxidising agent **CO₂**
- d
- c

23. 4.7 Neutralisation Reactions

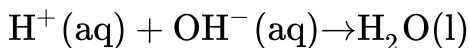
Practice questions

1. According to the Arrhenius definition, an acid is a compound that increases hydrogen ion concentration in an aqueous solution. True or false?
2. According to the Arrhenius definition, a base is a compound that increases hydroxyl ion concentration in an aqueous solution. True or false?
3. Predict the products of each acid-base combination listed. Assume that a neutralisation reaction occurs.



4. Write a balanced chemical equation for the neutralisation reaction between each given acid and base. Include the proper phase labels.
 - a. $\text{HI}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow ?$
 - b. $\text{H}_2\text{SO}_4(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow ?$
5. What is the balanced equation for the neutralisation reaction between $\text{HBr}(\text{aq})$ and $\text{Ba}(\text{OH})_2(\text{aq})$?
 - a. $2\text{HBr}(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{BaBr}_2(\text{aq})$
 - b. $2\text{HBr}(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{BaBr}_2(\text{aq})$

- c. $2\text{HBr}(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{Ba}(\text{OH})_2(\text{aq})$
 $\text{2HBr}(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq})$
- d. $\rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{BaBr}_2(\text{aq})$
6. Write the net ionic equation for $\text{HI}(\text{aq})$ and $\text{KOH}(\text{aq})$.
- a. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$
- b. $\text{HI}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{KI}(\text{s})$
- c. $\text{H}^+(\text{aq}) + \text{I}^-(\text{aq}) + \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{KI}(\text{s})$
- d. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) + \text{K}^+(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{I}^-$
7. Write the balanced net ionic equation for neutralisation reaction between $\text{H}_2\text{SO}_4(\text{aq})$ and $\text{Ba}(\text{OH})_2(\text{aq})$.
- a. $\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{H}_2\text{O}(\text{l})$
- b. $2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Ba}(\text{s}) + \text{SO}_4^{2-}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
- c. $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 $\rightarrow \text{H}_2\text{O}(\text{l})$
 $2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
- d. $\text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$
 $\rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
8. Why is the net ionic equation for the neutralisation reaction between $\text{HCl}(\text{aq})$ and $\text{KOH}(\text{aq})$ the same as the net ionic equation for the neutralisation reaction between $\text{HNO}_3(\text{aq})$ and RbOH ?
- a. Because both reactions are acid-base neutralisation reactions.
- b. Because the salts are insoluble/solids in both cases.
- c. Because the salts are soluble in both cases, the net ionic reaction is just



- d. Because H^+ and OH^- are spectator ions in both cases.
9. Write a balanced chemical equation for the neutralisation reaction between HNO_3 and $\text{Fe}(\text{OH})_3$.
- $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$
 - $3\text{HNO}_3(\text{aq}) + \text{Fe}(\text{OH})_3(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{NO}_3^-(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
 - $3\text{HNO}_3(\text{aq}) + \text{Fe}(\text{OH})_3(\text{aq}) \rightarrow \text{Fe}(\text{NO}_3)_3(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
 - $\text{HNO}_3(\text{aq}) + \text{Fe}(\text{OH})_3(\text{aq}) \rightarrow \text{Fe}(\text{NO}_3)_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
10. Which of the following is a neutralisation reaction?
- $2\text{NaHCO}_3(\text{aq}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - $\text{Al}(\text{OH})_3(\text{aq}) + 3\text{HNO}_3(\text{aq}) \rightarrow \text{Al}(\text{NO}_3)_3(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
 - $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
 - $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$

Answers

- True
- True
- Products are as follows:
 - $\text{KCl} + \text{H}_2\text{O}$
 - $\text{K}_2\text{SO}_4 + \text{H}_2\text{O}$
 - $\text{Ni}_3(\text{PO}_4)_2$
- Balanced chemical equations are as follows:
 - $\text{HCl} + \text{KOH} \rightarrow \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - $\text{H}_2\text{SO}_4(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
- a

- 6. a
- 7. d
- 8. c
- 9. c
- 10. b

24. 5.1 The Mole and Avogadro's Number

Practice questions

Short-answer questions

1. Determine the number of particles for the following number of moles. Round to 2 decimal places.
 - a. 1 mol= _____ particles
 - b. 2 mol= _____ particles
 - c. 5 mol= _____ particles
 - d. 0.5 mol= _____ particles
 - e. 0.1 mol= _____ particles
2. Determine the number of moles for the following number of particles.
 - a. 6.02×10^{23} = _____ mol
 - b. 1.204×10^{23} = _____ mol
 - c. 1.4448×10^{25} = _____ mol
 - d. 2.2274×10^{23} = _____ mol
 - e. 1.6555×10^{24} = _____ mol

Solutions

1.
 - a. 1 mol= 6.02×10^{23} particles
 - b. 2 mol= 1.20×10^{24} particles
 - c. 5 mol= 3.01×10^{24} particles

- d. $0.5 \text{ mol} = 3.01 \times 10^{23} \text{ particles}$
 - e. $0.1 \text{ mol} = 3.02 \times 10^{22} \text{ particles}$
- 2.
- a. $6.02 \times 10^{23} = 1 \text{ mol}$
 - b. $1.204 \times 10^{23} = 0.2 \text{ mol}$
 - c. $1.4448 \times 10^{25} = 24 \text{ mol}$
 - d. $2.2274 \times 10^{23} = 0.37 \text{ mol}$
 - e. $1.6555 \times 10^{24} = 2.75 \text{ mol}$

25. 5.2 Molar Mass

Practice questions

Multiple choice questions

- How many acetaminophen molecules are present in 0.25mol of acetaminophen?
 - 24.088×10^{23} acetaminophen molecules
 - 1.5×10^{23} acetaminophen molecules
 - 0.042×10^{23} acetaminophen molecules
 - 6.022×10^{23} acetaminophen molecules
- Glucose has the formula $C_6H_{12}O_6$. Calculate the molar mass of glucose. The molar masses of C, H, and O are given here: C = 12.01g/mol, H = 1.01g/mol, O = 16.00g/mol
 - 174.12g/mol
 - 29.02g/mol
 - 53.02g/mol
 - 180.1g/mol
- How many moles of C are present in 1.5mol of glucose? The chemical formula of glucose is $C_6H_{12}O_6$.
 - 9mol
 - 18mol
 - 4mol
 - 0.25mol
- How many moles of H are present in 1.5mol of glucose? The chemical formula of glucose is $C_6H_{12}O_6$.
 - 9mol
 - 18mol

- c. 4mol
d. 0.25mol
5. How many moles of O are present in 1.5mol of glucose? The chemical formula of glucose is $C_6H_{12}O_6$.
- a. 9mol
b. 18mol
c. 4mol
d. 0.25mol

Short-answer questions

1. Calculate the amount of molecular oxygen O_2 present in 160.0g of oxygen. The molar mass of atomic O is 16.00g/mol.

Step 1: Calculate the molar mass of molecular oxygen:

_____ g/mol \times _____ = _____ g/mol

Step 2: Calculate the moles of molecular oxygen (HINT: Molar Mass = Mass \div Moles): Moles n

= _____ \div _____ = _____ mol

2. Calculate how many grams of carbon are present in 1.8g of glucose.

The chemical formula of glucose is $C_6H_{12}O_6$. The atomic molar masses of C, H and O are given:

C = 12.01g/mol, H = 1.01g/mol, O = 16.00g/mol. Please round the answers to two d.p.

- a. Step 1 – Calculate the molar mass of glucose 12.01g/mol \times

_____ + 1.01g/mol \times _____ + 16.00g/mol \times _____ = _____ g/mol

- b. Step 2 – Calculate the number of moles of glucose present in 1.8g of glucose $n = m/M$

$n =$ _____ \div 180.1 = _____ mol

- c. Step 3 – Calculate the number of moles of carbon present

in the number of moles of glucose from step 2. 6mol C in
 $1\text{mol C}_6\text{H}_{12}\text{O}_6 \times \text{_____ mol C}_6\text{H}_{12}\text{O}_6 =$
 _____ mol C

- d. Step 4 – Convert the number of moles of carbon into grams $m = n \times M$
 $m = \text{_____} \times 12.01\text{g/mol}$
 $m = \text{_____ g}$

Solutions

Multiple choice questions

- b
- d Glucose contains six atoms of C, twelve atoms of H and six atoms of O. $(6 \times 12.01\text{g/mol}) + (12 \times 1.01\text{g/mol}) + (6 \times 16.00\text{g/mol}) = 180.1\text{g/mol}$.
- a
- b
- a

Short-answer questions

- Step 1: $16\text{g/mol} \times 2 = 32\text{g/mol}$
 Step 2: Moles $n = 160 \div 32 = 5\text{mol}$
- Step 1: $12.01\text{g/mol} \times 6 + 1.01\text{g/mol} \times 12 + 16.00\text{g/mol} \times 6 = 180.10\text{g/mol}$
 Step 2: $n = 1.8 \div 180.1 = 0.01\text{mol}$
 Step 3: 6mol C in 1mol $\text{C}_6\text{H}_{12}\text{O}_6 \times 0.01\text{mol C}_6\text{H}_{12}\text{O}_6 = 0.06\text{mol C}$
 Step 4: $m = n \times M = 0.06 \times 12.01\text{g/mol} = 0.72\text{ g}$

26. 5.3 Concentration and Solutions

Figures

Figure 5.3.3

A 3D molecular model displaying a salt crystal lattice interspersed with water and potassium molecules, as an illustration of potassium chloride dissolving into water. There are spheres labeled K^+ , representing potassium ions, and other spheres labeled Cl^- , representing chloride ions, arranged in an alternating cubic pattern. Surrounding this ionic lattice are several water molecules, and the orientation of the water molecules suggests hydration shells, where the oxygen atoms are attracted to the potassium ions and the hydrogen atoms are oriented towards the chloride ions, highlighting the ionic dissolution process in water.

Practice questions

Multiple choice questions

1. A sodium chloride solution contains 3.56g of sodium chloride dissolved in a final volume of 25.0 mL. What is the percent concentration of sodium chloride in m/v?
 - a. 14.2% m/v

- b. 7.02% m/v
 - c. 89% m/v
 - d. 85.5% m/v
2. When 2.50g of LiBr is dissolved in 25.0g of water, it gives a 10.0% m/m LiBr solution. Calculate the mass of LiBr present in 15.0g of the same solution.
- a. 0.150g
 - b. 1.50g
 - c. 1g
 - d. 10g
3. A diluted NaOH solution was prepared by adding 100.0mL of water to the 400.0mL of the 0.1M stock solution. What is the molarity of the diluted NaOH solution?
- a. 0.08M
 - b. 25M
 - c. 0.04M
 - d. 12.5M
4. 50.0mL of a NaOH stock solution with a concentration of 10.0% m/v is diluted to 250.0mL. What is the final concentration of the NaOH solution in m/v %?
- a. 5.00% m/v
 - b. 0.02% m/v
 - c. 2.00% m/v
 - d. 20% m/v

Short-answer questions

1. Calculate the molarity of the following solutions, then put the correct answer into the boxes. Round the answers to 2 d.p. if needed.
- 3.0mol of Silver nitrate in 0.50L of solution- _____

M

- 0.0625mol of Sodium bicarbonate in 250.0mL of solution- _____M
2. Calculate the mass in grams of the solute in each of the following solutions, then put the correct answer into the spaces. Round the answers to 2 d.p. if needed.
- 750.0mL of a 1.50M NaOH solution- _____ g
 - 0.200L of a 150.0mM NaCl solution- _____ g

Note: The molar masses of NaOH and NaCl are 39.997g/mol and 58.44g/mol, respectively.

3. Calculate the molarity of a 5.0%*m/v* KOH solution, then put the correct answer into the space.

Note: the molar mass of KOH is 56.1056g/mol. Round the answers to 2 d.p. if needed.

The Molarity of a 5.0% *m/v* KOH solution= _____ M

Solutions

Multiple choice questions

1. a
2. b
3. a
4. c

Short-answer questions

1. 6; 0.25

2. 45; 1.75

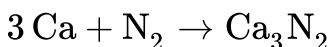
3. 0.89

27. 5.4 Limiting Reagents and Yield

Practice questions

Short-answer questions

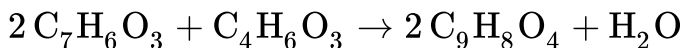
1. Calcium nitride Ca_3N_2 is produced using Ca and N_2 as follows:



What is the theoretical yield of Ca_3N_2 when 120.56g of Ca is reacted with 42.87g of N_2 ? The atomic molar mass of Ca is 40.08g/mol, and the atomic molar mass of N is 14.01g/mol. The molar mass of Ca_3N_2 is 148.26g/mol.

Answer: _____ g. The answers should be rounded to 2 d.p.

2. Salicylic acid $\text{C}_7\text{H}_6\text{O}_3$ reacts with acetic anhydride $\text{C}_4\text{H}_6\text{O}_3$ in excess to form acetylsalicylic acid $\text{C}_9\text{H}_8\text{O}_4$, which we call aspirin. The percentage yield of aspirin from the following reaction is 88.48%.



How many grams of aspirin were actually formed by the reaction of 250.82g of salicylic acid? The molar masses of salicylic acid and acetylsalicylate acid are 138.21g/mol and 180.16g/mol, respectively.

Answer: _____ g. The answers should be rounded to 2 d.p.

Solutions

1. 148.75 g
2. 289.30 g

28. 5.5 Standard Lab Conditions and the Ideal Gas Law

Practice questions

Multiple choice questions

1. In Boyle's law, which of the following remains constant?
 - a. Pressure
 - b. Temperature
 - c. Volume
 - d. Both temperature and volume
2. Which of the following statements is false?
 - a. The volume of a gas increases when the temperature increases at constant pressure.
 - b. Avogadro's Law applies at constant pressure.
 - c. Charles' Law applies at constant temperature.
 - d. An ideal gas obeys the gas laws.

True/false questions

1. Scientists have to record in their writings or laboratory reports changes in factors such as temperature, pressure, humidity, or any other external factors that cause any major effects on the outcome of an experiment.

Solutions

Multiple choice questions

1. b
2. c

True/false question

1. True

29. 6.1 Definitions and Theories of Acids and Bases

Practice questions

1. Identify each substance as an Arrhenius acid, an Arrhenius base, or neither.
 - a. NaOH
 - b. $\text{C}_2\text{H}_5\text{OH}$
 - c. H_3PO_4
 - d. $\text{Ba}(\text{OH})_2$
 - e. HNO_2
 - f. $\text{C}_6\text{H}_{12}\text{O}_6$
2. What is the balanced chemical equation for the neutralisation reaction between KOH and $\text{H}_2\text{C}_2\text{O}_4$?
 - a. $\text{KOH} + \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{H}_2\text{O} + \text{K}_2\text{C}_2\text{O}_4$
 - b. $2\text{KOH} + \text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{K}_2\text{C}_2\text{O}_4$
 - c. $2\text{KOH} + \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{H}_2\text{O} + \text{H}_2\text{C}_2\text{O}_4$
 - d. $2\text{KOH} + \text{H}_2\text{C}_2\text{O}_4 \rightarrow 2\text{H}_2\text{O} + \text{K}_2\text{C}_2\text{O}_4$
3. Is the following statement true? The balanced chemical equation for the neutralisation reaction between HCl and $\text{Fe}(\text{OH})_3$ is $3\text{HCl} + \text{Fe}(\text{OH})_3 \rightarrow 3\text{H}_2\text{O} + \text{FeCl}_3$.

.
4. How does *Brønsted-Lowry acid* differ from an Arrhenius acid?
 - a. A Brønsted-Lowry acid is a proton

- donor. It decreases the H^+ concentration in water.
- A Brønsted-Lowry acid is a proton donor. It increases the OH^- concentration in water.
 - A Brønsted-Lowry acid is a proton donor. It does not necessarily increase the H^+ concentration in water.
 - A Brønsted-Lowry acid is a proton donor. It increases the H^+ in water.
- The following equation shows the dissociation of hydrogen bromide in water as a Brønsted-Lowry acid-base reaction. Identify the proton donor and proton acceptor in this reaction:

$$\text{HBr} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Br}^-.$$
 - Pyridine ($\text{C}_5\text{H}_5\text{N}$) acts as a Brønsted-Lowry base in water. The hydrolysis reaction for pyridine is shown below. Identify the Brønsted-Lowry acid and Brønsted-Lowry base in this reaction:

$$\text{C}_5\text{H}_5\text{N} + \text{H}_2\text{O} \rightarrow \text{C}_5\text{H}_5\text{NH}^+ + \text{OH}^-.$$
 - Identify the Brønsted-Lowry acid and Brønsted-Lowry base in this chemical equation:

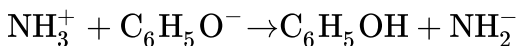
$$\text{H}_3\text{PO}_4 + \text{OH}^- \rightarrow \text{H}_2\text{PO}_4^- + \text{H}_2\text{O}.$$
 - Predict the products of this reaction, assuming it undergoes a Brønsted-Lowry acid-base reaction:

$$\text{HC}_2\text{H}_3\text{O}_2 + \text{C}_5\text{H}_5\text{N} \rightarrow ?$$

- a. $\text{C}_2\text{H}_3\text{O}_2^-$ and $\text{C}_5\text{H}_5\text{NH}^+$
 - b. $\text{C}_2\text{H}_3\text{O}_2^+$ and $\text{C}_5\text{H}_5\text{NH}^-$
 - c. H_2O and $\text{C}_5\text{H}_5\text{NH}^+$
 - d. $\text{C}_2\text{H}_3\text{O}_2^-$ and H_2O
9. What is the conjugate acid of: (A) H_2O and (B) NH_3 ?
 - a. $\text{A} = \text{H}_3\text{O}^+$, $\text{B} = \text{NH}_4^+$
 - b. $\text{A} = \text{CH}_3\text{OH}$, $\text{B} = \text{N}_2$
 - c. $\text{A} = \text{H}_2\text{O}^+$, $\text{B} = \text{NH}_3^+$
 - d. $\text{A} = \text{OH}^-$, $\text{B} = \text{NH}_2^-$
10. What is the conjugate base of: (A) HSO_4^- and (B) H_2O ?
 - a. $\text{A} = \text{S}$, $\text{B} = \text{H}^+$
 - b. $\text{A} = \text{SO}_4^{2-}$, $\text{B} = \text{H}^+$
 - c. $\text{A} = \text{HSO}_4^{2-}$, $\text{B} = \text{OH}^-$
 - d. $\text{A} = \text{SO}_4^{2-}$, $\text{B} = \text{OH}^-$
11. Identify the conjugate acid-base pairs in this reaction:

$$\text{HSO}_4^- + \text{PO}_4^{3-} \rightarrow \text{SO}_4^{2-} + \text{HPO}_4^{2-}$$
 - a. HSO_4^- and SO_4^{2-} ; PO_4^{3-} and HPO_4^{2-} .
 - b. HSO_4^- and PO_4^{3-} ; SO_4^{2-} and HPO_4^{2-} .
 - c. HSO_4^- and HPO_4^{2-} ; SO_4^{2-} and PO_4^{3-} .
 - d. None of the above

12. Identify the conjugate acid-base pairs in this reaction:



- a. NH_3 and NH_2^- ; $\text{C}_6\text{H}_5\text{O}^-$ and $\text{C}_6\text{H}_5\text{OH}$
- b. NH_3 and $\text{C}_6\text{H}_5\text{O}^-$; $\text{C}_6\text{H}_5\text{OH}$ and NH_2^-
- c. NH_3 and $\text{C}_6\text{H}_5\text{OH}$; NH_2^- and $\text{C}_6\text{H}_5\text{OH}^-$
- d. None of the above

Answers

1. (a) and (d) are Arrhenius bases, (c) and (e) are Arrhenius acids, (b) and (f) are neither
2. d
3. True
4. c
5. HBr – proton donor, H_2O – proton acceptor
6. H_2O – Brønsted-Lowry acid, $\text{C}_5\text{H}_5\text{N}$ – Brønsted-Lowry base
7. H_3PO_4 – Brønsted-Lowry acid, OH^- – Brønsted-Lowry base
8. a
9. a
10. d
11. a

12. a

30. 6.2 Autoionisation of Water

Practice questions

1. Does hydrogen ion concentration H^+ remain constant in all aqueous solutions?
 - a. No, hydrogen ion concentration varies with the amount of water present in an aqueous solution.
 - b. Yes, hydrogen ion concentration remains constant in all aqueous solutions because of the amount of water.
 - c. Yes, because $\text{H}^+ \times \text{OH}^- = 1 \cdot 0 \times 10^{-14}$
 - d. No, hydrogen ion concentration varies with the amount of acid or base present in an aqueous solution.
2. For a given aqueous solution, if $[\text{H}^+] = 1 \cdot 0 \times 10^{-3} \text{ M}$, what is $[\text{OH}^-]$?
 - a. $1 \cdot 0 \times 10^{-11}$
 - b. $1 \cdot 0 \times 10^{-6}$
 - c. $1 \cdot 0 \times 10^{-17}$
 - d. $1 \cdot 0 \times 10^{-14}$
3. For a given aqueous solution, if $[\text{H}^+] = 7 \cdot 92 \times 10^{-5} \text{ M}$, what is $[\text{OH}^-]$?
 - a. $0 \cdot 89 \times 10^{-10} \text{ M}$
 - b. $1 \cdot 26 \times 10^{-10} \text{ M}$
 - c. $1 \cdot 26 \times 10^{-9} \text{ M}$

- d. $1 \cdot 26 \times 10^{-19} \text{M}$
4. For a given aqueous solution, if $[\text{OH}^-] = 1 \cdot 0 \times 10^{-5} \text{M}$, what is H^+ ?
- a. $5 \cdot 0 \times 10^{-9} \text{M}$
 b. $1 \cdot 0 \times 10^{-6} \text{M}$
 c. $1 \cdot 0 \times 10^{-9} \text{M}$
 d. $2 \cdot 5 \times 10^{-9} \text{M}$
 e. $1 \cdot 0 \times 10^{-14} \text{M}$
5. For a given aqueous solution, if $[\text{OH}^-] = 3 \cdot 77 \times 10^{-4} \text{M}$, what is H^+ ?
- a. $2 \cdot 65 \times 10^{-10} \text{M}$
 b. $0 \cdot 65 \times 10^{-12} \text{M}$
 c. $2 \cdot 65 \times 10^{-11} \text{M}$
 d. $1 \cdot 0 \times 10^{-11} \text{M}$
6. What are $[\text{H}^+]$ and $[\text{OH}^-]$ in a $0 \cdot 344 \text{M}$ solution of HNO_3 ?
- a. $[\text{H}^+] = 0 \cdot 657 \text{M}$ and $[\text{OH}^-] = 0 \cdot 91 \times 10^{-10} \text{M}$
 b. $[\text{H}^+] = 1 \cdot 0 \times 10^{-10} \text{M}$ and $[\text{OH}^-] = 1 \cdot 0 \times 10^{-14} \text{M}$
 c. $[\text{H}^+] = 5 \cdot 0 \times 10^{-12} \text{M}$ and $[\text{OH}^-] = 2 \cdot 5 \times 10^{-14} \text{M}$
 d. $[\text{H}^+] = 0 \cdot 344 \text{M}$ and $[\text{OH}^-] = 2 \cdot 91 \times 10^{-14} \text{M}$
7. What are $[\text{H}^+]$ and $[\text{OH}^-]$ in a $0 \cdot 00338 \text{M}$ solution of KOH ?
- a. $[\text{H}^+] = 2 \cdot 96 \times 10^{-12} \text{M}$ and

$$[\text{OH}^-] = 0.00338 \text{ M}$$

b. $[\text{H}^+] = 0.00338 \text{ M}$ and

$$[\text{OH}^-] = 1.56 \times 10^{-12} \text{ M}$$

c. $[\text{H}^+] = 0.96 \times 10^{-12} \text{ M}$ and

$$[\text{OH}^-] = 1 \times 10^{-8} \text{ M}$$

d. $[\text{H}^+] = 0.96 \times 10^{-12} \text{ M}$ and

$$[\text{OH}^-] = 1 \times 10^{-10} \text{ M}$$

8. If HNO_2 is dissociated only to an extent of 0.445, what are $[\text{H}^+]$ and $[\text{OH}^-]$ in a 0.307 M solution of HNO_2 ?

a. $[\text{H}^+] = 0.00137 \text{ M}$ and

$$[\text{OH}^-] = 0.78 \times 10^{-10} \text{ M}$$

b. $[\text{H}^+] = 0.137 \text{ M}$ and

$$[\text{OH}^-] = 0.32 \times 10^{-10} \text{ M}$$

c. $[\text{H}^+] = 1.37 \times 10^{-12} \text{ M}$ and

$$[\text{OH}^-] = 7.32 \times 10^{-12} \text{ M}$$

d. $[\text{H}^+] = 0.00137 \text{ M}$ and

$$[\text{OH}^-] = 7.32 \times 10^{-12} \text{ M}$$

Answers

1. d
2. a
3. b
4. c

5. c
6. d
7. a
8. d

3I. 6.3 The pH Scale

Practice questions

- What is pH?
 - pH is the negative logarithm of $[\text{OH}^-]$
 - pH is the negative logarithm of $[\text{H}^+]$
 - pH is the logarithm of $[\text{H}^+]$
 - pH is the logarithm of $[\text{OH}^-]$
- What is the pH range for an acidic solution?
 - pH=7
 - pH < 7
 - pH > 7
 - pH=14
- What is the pH range for a basic solution?
 - pH=7
 - pH < 7
 - pH > 7
 - pH=1
- What is $[\text{H}^+]$?
 - $1.0 \times 10^{-7} \text{ M}$
 - $1.0 \times 10^{-14} \text{ M}$
 - $1.0 \times 10^{-1} \text{ M}$
 - $1.0 \times 10^{-8} \text{ M}$
- What is the pH of a solution when $[\text{H}^+]$ is $3.44 \times 10^{-4} \text{ M}$?
 - pH of 3.6

- b. $1.0 \times 10^{-4} M$
c. pH of 4
d. pH of 8
6. What is the pH of a solution when $[\text{OH}^-]$ is $6.22 \times 10^{-7} M$?
a. pH of 0.79
b. pH of 2
c. pH of 7.79
d. pH of 6
7. What is the pOH of a solution when $[\text{H}^+]$ is $3.44 \times 10^{-4} M$?
a. pOH of 1.89
b. pOH of 10.54
c. pOH of 3.46
d. pOH of 0.54
8. What is the pOH of a solution when $[\text{OH}^-]$ is $6.22 \times 10^{-7} M$?
a. pOH of 3.46
b. pOH of 7.79
c. pOH of 6.21
d. pOH of 7
9. If a solution has a pH of **0.77**, what is its pOH, $[\text{H}^+]$, and $[\text{OH}^-]$?
10. If a solution has a pOH of **13.9**, what is its pH, $[\text{H}^+]$, and $[\text{OH}^-]$?

Answers

1. b
2. b

3. c
4. a
5. a
6. c
7. b
8. c
9. $pOH = 13.23$, $[H^+] = 1 \cdot 70 \times 10^{-1} M$,
 $[OH^-] = 5 \cdot 89 \times 10^{-14}$
10. $pH = 0.91$, $[H^+] = 8 \cdot 13 \times 10^{-14} M$,
 $[OH^-] = 0 \cdot 12 M$

32. 6.4 Strength of Acids and Bases and their Salts

Practice questions

1. Is this statement true? A strong acid completely dissociates (100% ionised) in an aqueous solution, whereas a weak acid does not completely dissociate.
2. Identify each as a strong acid or a weak acid. Assume aqueous solutions.
 - a. HF
 - b. HCl
 - c. $\text{H}_2\text{C}_2\text{O}_4$
3. Identify each as a strong base or a weak base. Assume aqueous solutions.
 - a. NaOH
 - b. $\text{Al}(\text{OH})_3$
 - c. $\text{C}_4\text{H}_9\text{NH}_2$
4. Write a chemical equation for the ionisation of each following acid and indicate whether it proceeds 100% to products or not.
 - a. HNO_3
 - b. HNO_2
 - c. HI_3
5. Identify each salt as neutral, acidic, or basic.
 - a. NaNO_2
 - b. NaNO_3
 - c. NH_4NO_3

Answers

1. True
2. (a) weak acid, (b) strong acid, (c) weak acid
3. (a) strong base, (b) weak base, (c) weak base
4.
 - a. $\text{HNO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$ - proceeds 100%
 - b. $\text{HNO}_2(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_2^-(\text{aq})$ - does not proceed to 100%
 - c. $\text{HI}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{I}_3^-(\text{aq})$ - does not proceed to 100%
5.
 - a. basic
 - b. neutral
 - c. acidic

33. 6.5 Acid-Base Titrations

Practice questions

- Which of the following statements is incorrect regarding titration?
 - The titrant is usually added to the analyte, which is a substance with an unknown concentration.
 - Usually, the analyte (substance with an unknown concentration) is added to the titrant.
 - Titration is a chemical reaction performed in a quantitative fashion.
 - The reagent with the known concentration is typically the titrant.
- Is this statement true? An acid is always the titrant.
- If 55.60 mL of 0.2221 M HCl was needed to titrate a sample of NaOH to its equivalence point, what mass of NaOH was present?
 - 0.494 g
 - 8.76 g
 - 3.87 g
 - 56.34 g
- It takes 45.66 mL of 0.1126 M HBr to titrate 25.00 mL of $\text{Ca}(\text{OH})_2$ to its equivalence point. What is the original concentration of the $\text{Ca}(\text{OH})_2$ solution?
 - 2.897 M
 - 45.981 M
 - 10.378 M
 - 0.103 M

Answers

1. b
2. False
3. a
4. d

34. 6.6 Buffers

Practice questions

1. Is this statement true? A buffer is the combination of a weak acid or base and a salt of that weak acid or base.
2. Which combinations of compounds can make a buffer? Assume aqueous solutions. *Select all possible answers.*
 - a. HCl and NaCl
 - b. NH_4NO_3 and NH_3
 - c. HNO_2 and NaNO_2
 - d. NH_4NO_3 and HNO_3
3. The complete phosphate buffer system is based on four substances: H_3PO_4 , H_2PO_4^- , HPO_4^{2-} , and PO_4^{3-} .
What different buffer solutions can be made from these substances?

- a. H_3PO_4 and H_2PO_4^-
- b. H_2PO_4^- and HPO_4^{2-}
- c. HPO_4^{2-} and PO_4^{3-}
- d. All of the above

Answers

1. True
2. b, c
3. d

35. 7.1 Forms of energy

Figures

Figure 7.1.1

The typical process of a coal power electrical station in a flow chart. From left to right: coal is fed from a large grey container called 'coal supply' onto a conveyor belt into a furnace where it's burned. The heat from the furnace turns water into steam rising through boiler tubes. The steam then drives a turbine connected to a generator. After passing through the turbine, the steam is cooled back into water by a condenser. Electricity produced by the generator is carried away by 'transmission lines' leading to a utility pole.

Practice questions

Multiple choice questions

1. Which of the following describes the allowed actions of energy and matter under the Law of Conservation?
 - a. Create
 - b. Transform
 - c. Reconfigure
 - d. Destroy
 - e. Transmute
2. A filament lightbulb is connected to a power source and

activated. In what order does energy transform?

- a. Electrical Energy → Thermal Energy → Light Energy
 - b. Light Energy → Thermal Energy → Electrical Energy
 - c. Thermal Energy → Light Energy → Electrical Energy
3. In a combustion reaction, what type of energy is transformed into thermal energy?
- a. Light Energy
 - b. Kinetic Energy
 - c. Chemical Energy
 - d. Nuclear Energy

Solutions

Multiple choice questions

- 1. b
- 2. a
- 3. c

36. 7.2 Thermochemistry essentials

Figures

Figure 7.2.5: The electromagnetic spectrum

A diagram of the electromagnetic spectrum, labeled 'VISIBLE SPECTRUM' in the center. The spectrum has a color gradient from violet at 400 nanometres and labelled high energy, to red at 700 nanometres, labelled low energy. Below, the spectrum ranges are indicated: gamma rays from 0.0001 to 0.01 nanometers, X-rays from 0.01 to 10 nanometers, ultraviolet light from 10 nanometers to 400 nanometers. The visible spectrum spans from 400 to 700 nanometers, followed by infrared from 700 nanometers to 1 millimeter, and radio waves extending from 1 millimeter to 100 meters, subdivided into radar, TV, FM, and AM frequencies.

Figure 7.2.9: Types of systems within thermochemistry

Three types of thermodynamic systems using flask illustrations.

The first flask is open and shows an upward arrow marked mass and two horizontal arrows marked heat, indicating both mass and energy can be exchanged with the surroundings.

The second flask, closed, is sealed with a stopper and has 'mass' arrows mass inside the bottle pushing outwards and two heat

arrows pushing inwards, signifying that only energy can be exchanged, not mass.

The third flask is an isolated system and has internal mass and heat arrows pushing outwards on the flask walls, illustrating that neither mass nor energy is exchanged with the surroundings.

Practice questions

True/false question

1. In chemistry, temperature can be thought of as the kinetic energy of molecules.

Short-answer question

1. Absolute Zero is the temperature at which molecules stop moving. Its value is _____ K.
2. A _____ system doesn't allow matter and energy to be exchanged with the environment. A _____ system allows for energy, but not matter to be exchanged with the environment. A _____ system allows for both energy and matter to be exchanged with the environment.

Multiple choice questions

1. In fan-forced ovens, what form of heat transfer is primarily occurring?
 - a. Conduction
 - b. Convection

- c. Radiation
 - d. Transmittance
2. What forms of heat transfer can occur over a vacuum?
- a. Conduction
 - b. Convection
 - c. Radiation
 - d. Transmittance

Solutions

True/false question

1. True

Short-answer questions

1. 0 K
2. isolated, closed, open

Multiple choice questions

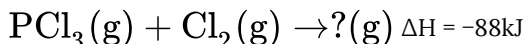
1. b
2. c

37. 7.3 Exothermic and Endothermic Reactions

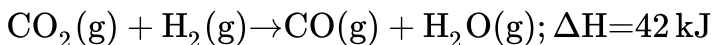
Practice questions

Multiple choice questions

1. Complete the thermochemical equation below for the reaction of $\text{PCl}_3(\text{g})$ with $\text{Cl}_2(\text{g})$ to make a $\text{PCl}_5(\text{g})$, which has an enthalpy change of -88kJ . Is this reaction exothermic or endothermic?



- a. $5\text{PCl}(\text{g})$; endothermic
 - b. $5\text{PCl}(\text{g})$; exothermic
 - c. $\text{PCl}_5(\text{g})$; endothermic
 - d. $\text{PCl}_5(\text{g})$; exothermic
2. Choose the correct thermochemical equation for the reaction of $\text{N}_2(\text{g})$ with $\text{O}_2(\text{g})$ to make $2\text{NO}(\text{g})$, which has an enthalpy change of 181kJ . Is this reaction exothermic or endothermic?
 - a. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$; exothermic
 - b. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$; endothermic
 - c. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2\text{O}_2(\text{g})$; exothermic
 - d. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{N}_2\text{O}_2(\text{g})$; endothermic
 3. Consider this thermochemical equation:

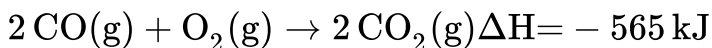


Is it exothermic or endothermic? How much energy is given off or absorbed?

- a. Endothermic; 42kj are absorbed
- b. Exothermic; 42kj are absorbed
- c. Endothermic; 42kj are given off
- d. Exothermic; 42kj are given off

True/false question

1. Consider this thermochemical equation:



Is the statement below is true or false?

This reaction is an exothermic reaction, and -565kj are released into the environment.

Solutions

Multiple choice questions

- 1. d
- 2. b
- 3. a

True/false question

1. False

38. 7.4 Calorimetry and Heat Capacity

Figures

Figure 7.4.1: Calorimetry

A diagram showing two stages of a calorimeter during chemical reactions. On the left (a), labeled 'System Exothermic process', a calorimeter contains a substance that, during an exothermic reaction, releases heat indicated by outward arrows, resulting in a rise in temperature shown by a red upward arrow on the thermometer. On the right (b), labeled 'System Endothermic process', it depicts an endothermic reaction where the substance absorbs heat, indicated by inward arrows, leading to a drop in temperature as shown by a red downward arrow on the thermometer. Both scenarios highlight the heat exchange between the system and the solution surrounding it.

Practice questions

Multiple choice questions (There is/ are one or more correct answer(s) for each question)

1. Which of the following are commonly used as units of energy within science?

- a. Joules
 - b. Moles
 - c. Kelvins
 - d. Calories
 - e. Newton
2. What type of thermodynamic system is a “coffee-cup” calorimeter?
- a. open
 - b. closed
 - c. isolated

True/false question

1. Endothermic reactions can be measured with a bomb calorimeter.

Short-answer question

1. The specific heat capacity of water is _____ $\text{JK}^{-1}\text{g}^{-1}$.

Solutions

Multiple choice questions

1. a, d
2. b

True/false question

1. False

Short-answer question

1. 4.18/ 4.186

39. 7.5 Rates of Reaction

Figures

Figure 7.5.3: Collision visualisations

A diagram of the concept of collision theory in chemical reactions. At the top is the general reaction $A + BX \text{ yields } B + AX$. Below, two sequences represent molecular interactions: The first sequence illustrates an ineffective collision where molecule A approaches molecule BX but does not result in a reaction, as indicated by the text ineffective collision text and an X mark. The second sequence shows an effective collision where molecule A approaches molecule BX from an orientation that allows a reaction to occur, leading to the formation of molecules B and AX, labeled with the green effective collision text and a check mark, signifying a successful reaction outcome.

Figure 7.5.4: Potential energy and activation energy

Graphical illustration of the course of a chemical reaction. The vertical axis is labeled potential energy indicating the energy of the molecules, while the horizontal axis is labeled reaction progress, representing the sequence of the reaction from start to finish.

The reactants start on the left with a certain level of energy, then the profile rises to a peak representing the activation energy barrier, labeled E_a . This is the minimum energy required for the reaction to proceed. After the peak the curve descends, with the energy of the

products lower than that of the reactants, suggesting the release of energy in the process.

Figure 7.5.3: Temperature and reaction rate

A graph illustration of the distribution of kinetic energy among molecules at two different temperatures.

The vertical axis is labeled fraction of molecules as the proportion of the total molecules at a given kinetic energy level. The horizontal axis is labeled kinetic energy showing the energy each molecule possesses due to motion.

Two bell-shaped curves are plotted: the taller bell curve peaks at a lower kinetic energy value signifying lower temperature; and the red curve, labeled higher temperature peaks at a higher kinetic energy value, showing that molecules at higher temperatures have a greater average kinetic energy.

The area under each curve to the right of the labeled activation energy, E_a , shaded with diagonal lines, represents the fraction of molecules with sufficient energy to overcome the activation energy barrier.

Practice questions

Short-answer question

1. Classify the following factors into the correct category.
Category A: Factors to increase the reaction rate and Category B: Factors to decrease the reaction rate:
 - Lower reactant concentration

- Higher reactant concentration
- Lower temperature
- Higher temperature
- A powdered reactant
- A solidified reactant
- Presence of a catalyst
- No catalyst present

True/false question

1. Increasing the temperature lowers the activation energy requirement.

Multiple choice question (There is/ are one or more correct answer(s) for each question)

1. What is required for a reaction to take place under collision theory?
 - a. Higher energy
 - b. Correct orientation
 - c. Sufficient energy
 - d. Presence of a catalyst

Solutions

Short-answer question

- Category A: Higher reactant concentration; Higher temperature; A powdered reactant; Presence of a catalyst
 - Category B: Lower reactant concentration; Lower temperature; A solidified reactant; No catalyst present

True/false question

1. False

Multiple choice question

1. b, c

40. 8.1 Introduction to Hydrocarbons

Practice questions

1. Which of the following are alkanes?
 - a. Cyclohexane ring: six carbons in a hexagon, C's are located at each point of the hexagon and connected by C-C single bonds.
 - b. Benzene ring: six carbons in a hexagon, C's are located at each point of the hexagon. Alternative C's form double bonds with the following C.
 - c. Two carbons are arranged in a continuous chain with one triple bond between two carbons.
 - d. Four carbons are arranged in a Y-shaped structure. Three CH₃ groups are located at the outer points of the Y-shape structure, while one CH₂ group situated at the junction of the Y-shape structure.
 - e. Eight carbons form a branching structure. The third and fifth carbons have methyl groups attached to those carbons.
2. Select all the alkene structures shown here:
 - a. Five carbons are located in a plus-shaped structure. Four outer carbons are methyl groups, while a carbon located at the middle point of the plus-shaped structure forms single bonds with four outer methyl groups.
 - b. Two carbon structures with a double bond between two carbons.
 - c. Five carbon structures with two C-C bonds: one

double bond between second and third carbon and the other between fourth and fifth carbon.

- d. Six carbon structures with a single, double bond in a branched chain. The second and third carbons have methyl groups attached to each carbon. $\text{C}-\text{C}$ bond located between second and third carbon.
 - e. Benzene ring: six carbons in a hexagon, C s are located at each point of the hexagon. Alternative C s form double bonds with the following C .
3. Which of the following belongs to the alkyne family?
- a. Two carbons are arranged in a continuous chain with one triple bond between two carbons.
 - b. A structure with one carbon, one hydrogen and one nitrogen. A triple bond is located between C and N . One hydrogen is attached to the carbon.
 - c. Eight carbons form a branched structure. Six carbons out of eight are arranged in a continuous chain with a $\text{C}-\text{C}$ bond located between the third and fourth carbons. The second and fifth carbons have methyl groups attached to each carbon.
 - d. Six carbons arranged in a continuous chain with two $\text{C}-\text{C}$ triple bonds located in the second and third carbons and the fourth and fifth carbons.
 - e. Six carbon structures with one double bond in a branched chain. The second and third carbons have methyl groups attached to each carbon. A $\text{C}-\text{C}$ bond is located between the second and third carbon.
4. Alkanes are less reactive compared to other hydrocarbons. Which of the following is the correct explanation for alkanes' reactivity?
- a. Alkanes are less reactive because of strong $\text{C}-\text{C}$ and $\text{C}-\text{H}$ single bonds.
 - b. Alkanes are less reactive because of the non-polar nature of $\text{C}-\text{C}$ and $\text{C}-\text{H}$ bonds.

- c. Alkanes are less reactive because they are polar molecules.
 - d. a and b
5. Which of the following belongs to the hydrocarbon family?
- a. A structure with two carbons. A carboxylic group (COOH) and a NH_2 attached to the first carbon.
 - b. Three benzene rings fused together.
 - c. One benzene ring with a $\text{CH}_2-\text{O}-\text{CH}=\text{O}$ attached to the first carbon of the benzene ring.
 - d. Five carbons located in a plus-shaped structure. Four outer carbons are methyl groups, while a carbon located at the middle point of the plus-shape structure forms single bonds with four outer methyl groups.
 - e. A four-carbon continuous chain. A $\text{C}-\text{C}$ triple bond is located between the second and the third carbons.
6. Explain the difference between saturated and unsaturated hydrocarbons by filling the blanks.

The difference between saturated and unsaturated hydrocarbons lies in the nature of the bonds: saturated hydrocarbons have only, while unsaturated hydrocarbons have at least one or bond. The carbon atoms in saturated hydrocarbons are saturated with atoms, meaning they are bonded to as many atoms as Unsaturated hydrocarbons have at least one carbon-carbon or bond, resulting in fewer atoms bonded to the atoms. This distinction affects the overall,, and properties of these hydrocarbons.

Answers

1. a, d, e
2. b, c, d, e
3. a, c, d
4. d
5. b, d, e
6. carbon-carbon, single, double, triple, hydrogen, hydrogen, possible, double, triple, hydrogen, carbon, structure, reactivity, physical

4I. 8.2 Drawing and Interpreting Organic Formulas

Practice questions

1. Determine how many total hydrogens are in each of these line-bond formulas.
 - a. A structure with three carbons. The first and third carbons each have one hydroxyl group attached to it. The second carbon has a carbon-oxygen double bond.
 - b. A structure with five carbons arranged in a straight chain. The first, fourth, and fifth carbons each have one hydroxyl attached. The first and the fifth carbons each have a carbon-oxygen double bond. The third carbon has a **COOH** group attached.
 - c. How many hydrogens for arginine are shown here? A structure with six-carbon in a straight chain, except **NH** group located between first and second carbons. The first and fifth carbons have **NH₂** attached. The first carbon is also bonded to an **NH** group by a double bond between nitrogen and carbon. The sixth carbon has a hydroxyl group and a double-bonded oxygen attached.
 - d. Two benzene rings fused.
2. How many carbon on this structure have no bonds to a hydrogen atom? The structure description: a benzene ring with a carboxylic group located at the first carbon. The second carbon has a **OC=OCH₃** attached.

- a. 1
 - b. 4
 - c. None
 - d. 3
3. How many hydrogens are bonded to carbon at each shown?
The structure description: a hexane ring with a four-carbon chain attached to the first carbon of the hexane ring. The four carbon chain has a carbon-carbon double bond located between the third and fourth carbons, a hydroxyl attached to the fourth carbon and an ethyl group attached to the first carbon. The points shown are the four carbons in the four-carbon chain and two carbons in the attached ethyl group.

Answers

- a. 6
 - b. 8
 - c. 14
 - d. 8
2. b
3. The first carbon on the four-carbon chain has one hydrogen. The second carbon on the four-carbon chain has two hydrogens. The third and the fourth carbons on the four-carbon chain have one hydrogen each. The first carbon in the ethyl group has two hydrogens, and the second carbon has three hydrogens.

42. 8.3 Naming Organic Compounds

Practice questions

1. Look up the IUPAC names for acetic acid, chloroform, and acetone. *One place you can find these is on Wikipedia, in the box of chemical information that is on the right side of the page. Wikipedia is quite reliable for this type of technical information. Other reliable sources for this kind of information include ChemSpider and PubChem.* What are the correct IUPAC names of the following chemicals: choose from these names—Trichloromethane, 2-propanone, and ethanoic acid.
 - a. Acetic acid,
 - b. Chloroform
 - c. Acetone
2. Which of these structures is 5-chloropentanal?
 - a. A five-carbon straight chain. The first carbon has a carbon-carbon double bond, and the fifth carbon has chlorine atoms attached.
 - b. A hexane ring with a methyl group attached to one of the carbons.
 - c. A hexane ring with two methyl groups attached to the second carbon and one oxygen attached to the first carbon forms a carbon-oxygen double bond.
3. Which structure shows 2,2-dimethylcyclohexanone?
 - a. A five-carbon straight chain. The first carbon has a carbon-carbon double bond, and the fifth carbon has chlorine atoms attached.
 - b. A hexane ring with a methyl group attached to one of

- the carbons.
- c. A hexane ring with two methyl groups attached to the second carbon and one oxygen attached to the first carbon forms a carbon-oxygen double bond.
4. Which of these is methylcyclohexane?
- a. A five-carbon straight chain. The first carbon has a carbon-carbon double bond, and the fifth carbon has chlorine atoms attached.
- b. A hexane ring with a methyl group attached to one of the carbons.
- c. A hexane ring with two methyl groups attached to the second carbon and one oxygen attached to the first carbon forms a carbon-oxygen double bond.
5. Which structure represents the substance named pentane?
- a. One carbon with four hydrogens attached.
- b. Five carbons in a straight chain with 12 hydrogens,
- c. Hexane ring.

Answers

1. a. ethanoic acid
b. trichloromethane
c. 2-propanone
2. a
3. c
4. b
5. b

43. 8.4 Naming and Drawing Hydrocarbons

Practice questions

1. Which name is correct for this structure? The structure description: five carbons with twelve hydrogens in a straight chain drawn in a line structure format.
 - a. Hexane
 - b. Pentane
 - c. Propane
 - d. Pentene
2. What name is best for this molecule? The structure description: five carbons with hydrogens in a straight chain drawn in a line structure format. The second and fourth carbons each have methyl groups.
 - a. 2,4-Dimethylpentane
 - b. 1,1-3,3-Tetramethylpropane
 - c. Heptane
 - d. 2,4-Dimethylheptane
3. What name is best for this molecule? The structure description: six carbons with hydrogens in a straight chain drawn in a line structure format. The second and fourth carbons each have methyl groups.
 - a. 3,5,5-Trimethylpentane
 - b. Dimethyloctane
 - c. 2,4-Dimethylhexane
 - d. 3,5-Dimethylhexane
4. Which name is best for this structure? The structure description: five carbons with a carbon-carbon triple bond

between the first and the second carbons drawn in a line structure format.

- a. 1-Pentene
 - b. Pent-1-yne
 - c. 1-Butyne
 - d. Pent-4-yne
5. What name is best for this structure? The structure description: four carbons with a carbon-carbon triple bond between the second and the third carbons drawn in a line structure format.
- a. Eth-2-yne
 - b. But-2-yne
 - c. Z-butyne
6. Choose the correct IUPAC names from the given list for the following structures: 2-ethyl-1-hexene, cyclohexane, cyclopentane, 5-methyl-1-hexene, 2,3-dimethyl-1-butene, 2-methyl-2-pentene, 3-ethyl-2-pentene.
- a. A structure with seven carbons in a straight chain and CH_2 group attached to the third carbon by a carbon-carbon double bond.
 - b. A five-carbon ring with a one-carbon double bond.
 - c. A structure with five carbons in a straight chain, a CH_3 group attached to the second carbon and a carbon double bond between the second and the third carbons.
 - d. A structure with four carbons in a straight chain, a CH_3 group attached to the second and third carbon each and a carbon double bond between the third and the fourth carbons.
 - e. Six carbon ring with one carbon double bond
 - f. A structure with six carbons in a straight chain, a CH_3 group attached to the fifth carbon and a carbon double bond between the first and the second

carbons.

- g. A structure with five carbons in a straight chain, a C_2H_5 group attached to the third carbon and a carbon double bond between the second and the third carbons.

7. Which of the following is the correct structure of 2-Methyl-1-pentene?

- a. A structure with five carbons in a straight chain, a CH_3 group attached to the second carbon and a carbon double bond between the first and the second carbons.
- b. A structure with five carbons in a straight chain, a CH_3 group attached to the fourth carbon and a carbon double bond between the second and the third carbons.
- c. A structure with five carbons in a straight chain, a CH_3 group attached to the fourth carbon and a carbon triple bond between the second and the third carbons.

8. What is the IUPAC name of this structure? The structure description: a structure with seven carbons in a straight chain, a C_2H_5 group attached to the fifth carbon and a carbon triple bond between the third and the fourth carbons.

- a. 5-methyl-2-heptyne
- b. 5-ethyl-2-heptene
- c. 5-ethyl-2-heptyne
- d. 3-ethyl-5-heptyne

9. What is the IUPAC name of this structure? The structure description: a structure with four carbons in a straight chain, a **Formula does not parse** group attached to the second carbon and a carbon double bond between the first and the second carbons.

- a. 2-propyl-1-propylene

- b. 2-Methyl-1-butane
 - c. 3-Methyl-3-butene
 - d. 2-Methyl-1-butene
10. What is the IUPAC name of this molecule? The structure description: a structure with three carbons in a straight chain and a carbon double bond between the first and the second carbons.
- a. 1-Propene
 - b. 2-Propene
 - c. 1-Propane
 - d. 2-Propyne

Answers

- 1. b
 - 2. c
 - 3. a
 - 4. b
 - 5. b
- 6.
- a. 2-ethyl-1-hexene,
 - b. cyclopentane,
 - c. 2-methyl-2-pentene,
 - d. 2,3-dimethyl-1-butene,
 - e. cyclohexane,
 - f. 5-methyl-1-hexene,
 - g. 3-ethyl-2-pentene.
- 7. a
 - 8. c
 - 9. d
 - 10. a

44. 8.5 Isomers

Practice questions

- Which structures among those shown here are related to one another as **constitutional isomers**? Select all that belong in the set.
 - A five-carbon ring with two carbon double bonds located in alternative carbons.
 - A four-carbon straight chain with a **NH** group located between the second and third carbon and a methyl group attached to the second and third carbons, each.
 - A five-carbon straight chain with two carbon double bonds: one double bond between the first and the second carbon and the other third and the fourth carbon.
 - A six-carbon straight chain with a **NH₂** attached to the last carbon.
- What is the constitutional isomer of ethanol, **CH₃CH₂OH**?
 - A two-carbon chain with a hydroxyl group attached to **CH₂**: **CH₃CH₂OH**
 - A two-carbon chain with a hydroxyl group attached to **CH₂**: **CH₃CH₂OH**
 - A two-carbon chain with an oxygen atom connecting two carbons: **CH₃OCH₃**
- Is it true that this is a constitutional isomer of resorcinol, molecular formula **C₆H₆O₂**? The structure description: a benzene ring with two hydroxyl groups attached to carbons

opposite to each other.

4. How many constitutional isomers does C_5H_{12} have?
- a. 3
 - b. None
 - c. 2
 - d. 4
5. How many constitutional isomers does $\text{C}_4\text{H}_{10}\text{O}$ have?
- a. 4
 - b. 3
 - c. 10
 - d. 7
6. How many structural isomers does $\text{C}_3\text{H}_9\text{N}$ have?
- a. 3
 - b. 2
 - c. 4
 - d. 7

Answers

- 1. b, d
- 2. c
- 3. True
- 4. a
- 5. 7
- 6. c

45. 8.6 Heteroatoms and Functional Groups

Practice questions

1. Identify which functional groups are present in the following structures.
 - a. A structure of a two-carbon chain with a **SH** group attached to the second carbon:
CH₃CH₂SH.
 - b. An eight-carbon chain with single carbon-carbon and carbon-hydrogen bonds and a carbon-oxygen double bond at the last carbon.
 - c. A three-carbon chain with a carbon double bond between the first and the second carbon.
 - d. Benzene ring with a **COOH** group attached to one of the carbons.
 - e. Hexane ring with a hydroxyl group attached to one of the carbons.
2. Select the structure of a secondary alcohol.
 - a. A three-carbon chain with a hydroxyl group is attached to the first carbon, and the second carbon has a methyl group attached.
 - b. A three-carbon chain with a hydroxyl group and two methyl groups attached to the first carbon.
 - c. A three-carbon chain with a hydroxyl group and a methyl group attached to the first carbon.
3. What are the functional groups present in this structure? If there are any alcohols present, identify those as primary, secondary or tertiary. The structure description: a six-carbon

straight chain with a double bonded oxygen and OCH_3 group attached to the first carbon. The fifth carbon has a methyl group and a hydroxyl attached.

4. What are the functional groups present in this structure? Select all the possible answers. The structure description: a hexagon with five carbons and one oxygen with one NH_2 group attached to the fourth position on the hexagon and $\text{CH}_2=\text{CH}_2$ group attached to the third position of the hexagon.
- a. An amine
 - b. An ether
 - c. An alkene
 - d. An Imine
5. Identify the following amines as primary, secondary or tertiary amines:
- a. One carbon with three hydrogens attached and one NH_2 attached.
 - b. One carbon with three hydrogens attached and one substituted nitrogen attached. Nitrogen is substituted with one hydrogen and one C_3H_7 .
 - c. One carbon with three hydrogens attached and one substituted nitrogen attached. Nitrogen is substituted with two methyl groups.
6. Which name describes a substance that contains the alcohol functional group?
- a. Cyclopentadiene
 - b. 3-chlorooctanol
 - c. 2-bromopentane
 - d. Dimethylamine
7. Which name describes a substance that contains the element nitrogen?
- a. 3-chlorooctanol
 - b. Cyclopentadiene

- c. 2-bromopentane
 - d. Dimethylamine
8. Which name describes a substance that is considered “unsaturated” (e.g. has a double or triple bond)?
- a. Cyclopentadiene
 - b. Dimethylamine
 - c. 3-chlorooctanol
 - d. 2-bromopentane
9. Which of these substances contain halogens?
- a. 3-chlorooctanol
 - b. Cyclopentadiene
 - c. Dimethylamine
 - d. 2-bromopentane
10. Which statement below is true?
- a. A given compound/molecule can contain one and only one functional group.
 - b. Compounds/molecules can and often do contain more than one functional group.

Answers

- 1.
 - a. thiol
 - b. aldehyde
 - c. alkene
 - d. carboxylic acid
 - e. alcohol
- 2. c
- 3. Ester and tertiary alcohol
- 4. a, b, c
- 5.
 - a. primary amine
 - b. secondary amine
 - c. tertiary amine

6. b
7. d
8. a
9. a, d
10. b

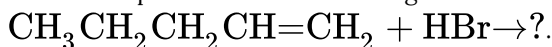
46. 8.7 Types of Organic Reactions

Practice questions

1. Classify the following organic reactions as addition, elimination and substitution.

- a. $\text{CH}_3\text{CH}_2\text{CH}=\text{CH}_2 + \text{H}_2 \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$
- b. $\text{CH}_3\text{CH}_2\text{CH}_2\text{Cl} + \text{KOH} \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{OH} + \text{KCl}$
- c. $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH} + \text{H}_2\text{SO}_4/\text{catalyst} \rightarrow \text{CH}_3\text{CH}=\text{CH}_2 + \text{H}_2\text{O}$

2. Predict the products of the following reaction:

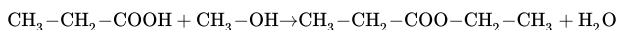


Select all the possible products.

- a. $\text{CH}_3\text{CHBr}-\text{CH}_2\text{CH}_2\text{CH}_3$
 - b. $\text{CHBr}-\text{CH}_2-\text{CH}_2-\text{CH}_2\text{CH}_3$
 - c. $\text{CH}_3-\text{CHBr}-\text{CHBr}-\text{CH}_2-\text{CH}_2-\text{CH}_3$
 - d. $\text{CH}_3\text{CHBr}-\text{CHBr}-\text{CH}_2\text{CH}_3$
3. Predict the reactants in the following reaction: the X reactant reacts with bromine to produce a chemical with the structure of a hexagon with 6 carbons 10 hydrogens, and two bromine atoms. Two bromine atoms are located at the first and second carbon of the hexagon: $\text{X} + \text{Br}_2 \rightarrow \text{C}_6\text{H}_{10}\text{Br}_2$.
- a. Hexagon with 6 carbons, 10 hydrogen and one carbon double bond.
 - b. Hexagon with 6 carbons, 8 hydrogen and two carbon double bonds.
 - c. Hexagon with 6 carbons, 6 hydrogen and

three carbon double bonds, each present in alternative carbons.

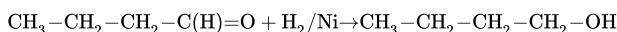
4. Classify the following organic reaction.



.

- a. Hydrolysis
- b. Addition
- c. Reduction
- d. Condensation

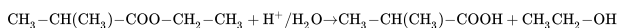
5. Classify the following organic reaction.



.

- a. Hydrolysis
- b. Reduction
- c. Condensation
- d. Oxidation

6. Classify the following organic reaction.



- a. Hydrolysis
- b. Addition
- c. Condensation
- d. Reduction

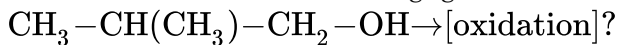
7. Predict the products of the following oxidation reaction of alcohols with an oxidising agent.



.

- a. $\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{C}(=\text{O})-\text{CH}_3$
- b. $\text{CH}_3-\text{CH}(\text{CH}_3)-\text{COOH}$
- c. $\text{CH}_3-\text{CH}_2-\text{C}(=\text{O})-\text{CH}_3$

8. Predict the products of the following oxidation reaction of alcohols with an oxidising agent.



.

- a. $\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{Br}$
 - b. $\text{CH}_3-\text{C}(\text{CH}_3)-\text{COOH}$
 - c. $\text{CH}_3-\text{CH}_2-\text{C}(=\text{O})-\text{CH}_3$
9. Predict the products of the following oxidation reaction of alcohols with an oxidising agent.
- $\text{CH}_3-\text{CH}_2-\text{CH}(\text{OH})-\text{CH}_3 \xrightarrow{\text{oxidation}} ?$
- a. $\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{C}(=\text{O})-\text{CH}_3$
 - b. $\text{CH}_3-\text{CH}(\text{CH}_3)-\text{COOH}$
 - c. $\text{CH}_3-\text{CH}_2-\text{C}(=\text{O})-\text{CH}_3$
10. Which of the following statements is correct regarding the products of the above reaction? The structure description: Hexagon with six carbons and 10 hydrogens, and one of the carbons is substituted with a hydroxyl group and a methyl group.
- $\text{C}_6\text{H}_{10}-\text{OH}(\text{CH}_3) \rightarrow [\text{oxidation}] ?$
- a. The above reaction produces a carboxylic acid.
 - b. The above reaction produces a ketone.
 - c. No reaction occurs.
 - d. The above reaction produces an aldehyde.

Answers

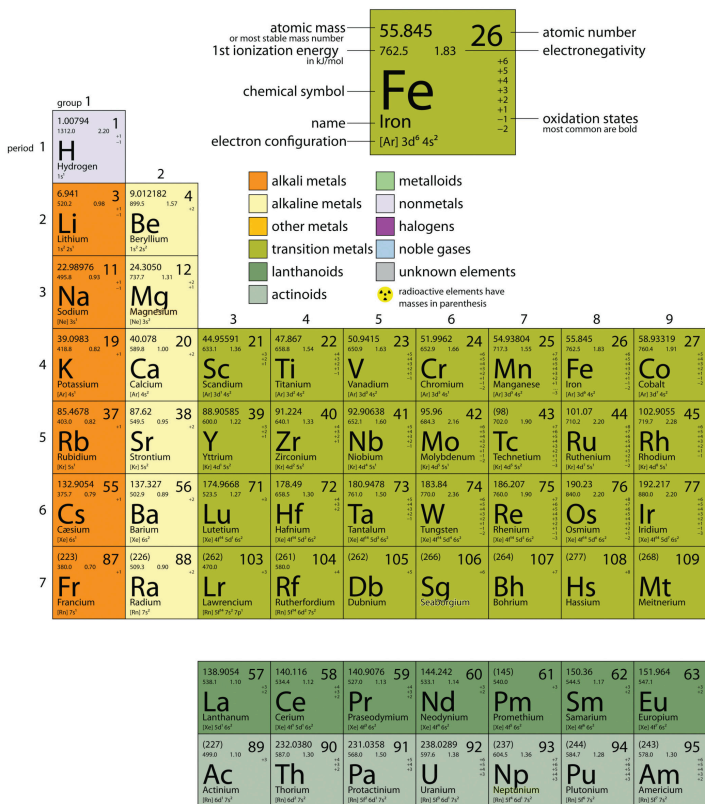
1.
 - a. Addition
 - b. Substitution
 - c. Elimination
2. a, b
3. a
4. d
5. b
6. a

7. a
8. b
9. c
10. c

In this appendix, we present some data on the chemical elements. The periodic table, introduced in Chapter 3 “Atoms, Molecules, and Ions,” lists all the known chemical elements, arranged by atomic number (that is, the number of protons in the nucleus). The periodic table is arguably the best tool in all of science; no other branch of science can summarise its fundamental constituents in such a concise and useful way. Many of the physical and chemical properties of the elements are either known or understood based on their positions on the periodic table. Periodic tables are available with a variety of chemical and physical properties listed in each element’s box. What follows here is a more complex version of the periodic table than what was presented in the chapter “Atoms, Molecules, and Ions.” The internet is a great place to find periodic tables that contain additional information.

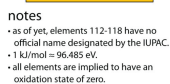
One item on most periodic tables is the atomic mass of each element. For many applications, only one or two decimal places are necessary for the atomic mass. However, some applications require more decimal places. The atomic masses in Table A.1 “The Basics of the Elements of the Periodic Table” represent the number of decimal places recognised by the International Union of Pure and Applied Chemistry, the worldwide body that develops standards for chemistry. The atomic masses of some elements are known very precisely, to a large number of decimal places. The atomic masses of other elements, especially radioactive elements, are not known as precisely. Some elements, such as lithium, can have varying atomic masses depending on how their isotopes are isolated.

The web offers many interactive periodic table resources. For example, see PTable and the Accessible Syngenta Periodic Table of Elements.



The properties listed in each box are introduced throughout the text. Atomic masses may vary by source.

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10	11	12	Aluminum (Al) 31 ^g	Silicon (Si) 28 ^g	Phosphorus (P) 31 ^g	Sulfur (S) 32 ^g	Chlorine (Cl) 35.5 ^g	Argon (Ar) 39.9 ^g
58.6934 28 Ni Nickel (Ni) 58.69 ^g	63.546 29 Cu Copper (Cu) 63.55 ^g	65.38 30 Zn Zinc (Zn) 65.38 ^g	69.723 31 Ga Gallium (Ga) 69.72 ^g	72.64 32 Ge Germanium (Ge) 72.64 ^g	74.9216 33 As Arsenic (As) 74.92 ^g	78.96 34 Se Selenium (Se) 78.96 ^g	79.904 35 Br Bromine (Br) 79.90 ^g	83.798 36 Kr Krypton (Kr) 83.80 ^g
106.42 46 Pd Palladium (Pd) 106.42 ^g	107.8662 47 Ag Silver (Ag) 107.87 ^g	112.4148 48 Cd Cadmium (Cd) 112.41 ^g	114.818 49 In Indium (In) 114.82 ^g	118.710 50 Sn Tin (Sn) 118.71 ^g	121.760 51 Sb Antimony (Sb) 121.76 ^g	127.603 52 Te Tellurium (Te) 127.60 ^g	126.9044 53 I Iodine (I) 126.90 ^g	131.293 54 Xe Xenon (Xe) 131.29 ^g
195.084 78 Pt Platinum (Pt) 195.08 ^g	196.9665 79 Au Gold (Au) 196.97 ^g	200.59 80 Hg Mercury (Hg) 200.59 ^g	204.3833 81 Tl Thallium (Tl) 204.38 ^g	207.2 82 Pb Lead (Pb) 207.2 ^g	208.9804 83 Bi Bismuth (Bi) 208.98 ^g	(210) 84 Po Polonium (Po) 210 ^g	(210) 85 At Astatine (At) 210 ^g	(220) 86 Rn Radon (Rn) 220 ^g
(271) 110	(272) 111	(285) 112	(284) 113	(289) 114	(288) 115	(292) 116	117	(294) 118
Ds Darmstadtium (Ds) 271 ^g	Rg Roentgenium (Rg) 272 ^g	Cn Copernicium (Cn) 285 ^g	Uut Ununtrium (Uut) 284 ^g	Uuq Ununquadium (Uuq) 289 ^g	Uup Unupentium (Uup) 288 ^g	Uuh Ununhexium (Uuh) 292 ^g	Uus Ununseptium (Uus) 292 ^g	Uuo Ununoctium (Uuo) 294 ^g

157.25 157.1	64 158.9233 Gd Gadolinium 64.6278	65 162.50 Tb Terbium 64.9290	66 164.9303 Dy Dysprosium 66.9298	67 167.259 Ho Holmium 67.9454	68 168.9342 Er Erbium 68.9359	69 173.054 Tm Thulium 69.9271	70 175.054 Yb Ytterbium 70.9254
181.0 180.9	96 247.0 Cm Curium 96.1059	97 251.0 Bk Berkelium 97.0145	98 259.1 Cf Californium 98.0792	99 262.1 Es Einsteinium 99.1071	100 267.1 Fm Fermium 100.1053	101 270.1 Md Mendelevium 101.0774	102 272.1 No Nobelium 102.0678

Table A.1 The Basics of the Elements of the Periodic Table.

Name	Atomic Symbol	Atomic Number	Atomic Mass	Footnotes
actinium*	Ac	89		
aluminum	Al	13	26.9815386(8)	
americium*	Am	95		
antimony	Sb	51	121.760(1)	g
argon	Ar	18	39.948(1)	g, r
arsenic	As	33	74.92160(2)	
astatine*	At	85		
barium	Ba	56	137.327(7)	
berkelium*	Bk	97		
beryllium	Be	4	9.012182(3)	
bismuth	Bi	83	208.98040(1)	
bohrium*	Bh	107		
boron	B	5	10.811(7)	g, m, r
bromine	Br	35	79.904(1)	
cadmium	Cd	48	112.411(8)	g
caesium (cesium)	Cs	55	132.9054519(2)	
calcium	Ca	20	40.078(4)	g
californium*	Cf	98		
carbon	C	6	12.0107(8)	g, r
cerium	Ce	58	140.116(1)	g
chlorine	Cl	17	35.453(2)	g, m, r
chromium	Cr	24	51.9961(6)	
cobalt	Co	27	58.933195(5)	
copernicium*	Cn	112		
copper	Cu	29	63.546(3)	r
curium*	Cm	96		
darmstadtium*	Ds	110		

Name	Atomic Symbol	Atomic Number	Atomic Mass	Footnotes
dubnium*	Db	105		
dysprosium	Dy	66	162.500(1)	g
einsteinium*	Es	99		
erbium	Er	68	167.259(3)	g
europium	Eu	63	151.964(1)	g
fermium*	Fm	100		
fluorine	F	9	18.9984032(5)	
francium*	Fr	87		
gadolinium	Gd	64	157.25(3)	g
gallium	Ga	31	69.723(1)	
germanium	Ge	32	72.64(1)	
gold	Au	79	196.966569(4)	
hafnium	Hf	72	178.49(2)	
hassium*	Hs	108		
helium	He	2	4.002602(2)	g, r
holmium	Ho	67	164.93032(2)	
hydrogen	H	1	1.00794(7)	g, m, r
indium	In	49	114.818(3)	
iodine	I	53	126.90447(3)	
iridium	Ir	77	192.217(3)	
iron	Fe	26	55.845(2)	
krypton	Kr	36	83.798(2)	g, m
lanthanum	La	57	138.90547(7)	g
lawrencium*	Lr	103		
lead	Pb	82	207.2(1)	g, r
lithium	Li	3	[6.941(2)]†	g, m, r
lutetium	Lu	71	174.967(1)	g
magnesium	Mg	12	24.3050(6)	
manganese	Mn	25	54.938045(5)	

Name	Atomic Symbol	Atomic Number	Atomic Mass	Footnotes
meitnerium*	Mt	109		
mendelevium*	Md	101		
mercury	Hg	80	200.59(2)	
molybdenum	Mo	42	95.94(2)	g
neodymium	Nd	60	144.242(3)	g
neon	Ne	10	20.1797(6)	g, m
neptunium*	Np	93		
nickel	Ni	28	58.6934(2)	
niobium	Nb	41	92.90638(2)	
nitrogen	N	7	14.0067(2)	g, r
nobelium*	No	102		
osmium	Os	76	190.23(3)	g
oxygen	O	8	15.9994(3)	g, r
palladium	Pd	46	106.42(1)	g
phosphorus	P	15	30.973762(2)	
platinum	Pt	78	195.084(9)	
plutonium*	Pu	94		
polonium*	Po	84		
potassium	K	19	39.0983(1)	
praseodymium	Pr	59	140.90765(2)	
promethium*	Pm	61		
protactinium*	Pa	91	231.03588(2)	
radium*	Ra	88		
radon*	Rn	86		
roentgenium*	Rg	111		
rhenium	Re	75	186.207(1)	
rhodium	Rh	45	102.90550(2)	
rubidium	Rb	37	85.4678(3)	g
ruthenium	Ru	44	101.07(2)	g

Name	Atomic Symbol	Atomic Number	Atomic Mass	Footnotes
rutherfordium*	Rf	104		
samarium	Sm	62	150.36(2)	g
scandium	Sc	21	44.955912(6)	
seaborgium*	Sg	106		
selenium	Se	34	78.96(3)	r
silicon	Si	14	28.0855(3)	r
silver	Ag	47	107.8682(2)	g
sodium	Na	11	22.98976928(2)	
strontium	Sr	38	87.62(1)	g, r
sulphur	S	16	32.065(5)	g, r
tantalum	Ta	73	180.94788(2)	
technetium*	Tc	43		
tellurium	Te	52	127.60(3)	g
terbium	Tb	65	158.92535(2)	
thallium	Tl	81	204.3833(2)	
thorium*	Th	90	232.03806(2)	g
thulium	Tm	69	168.93421(2)	
tin	Sn	50	118.710(7)	g
titanium	Ti	22	47.867(1)	
tungsten	W	74	183.84(1)	
ununhexium*	Uuh	116		
ununoctium*	Uuo	118		
ununpentium*	Uup	115		
ununquadium*	Uuq	114		
ununtrium*	Uut	113		
uranium*	U	92	238.02891(3)	g, m
vanadium	V	23	50.9415(1)	
xenon	Xe	54	131.293(6)	g, m
ytterbium	Yb	70	173.04(3)	g

Name	Atomic Symbol	Atomic Number	Atomic Mass	Footnotes
yttrium	Y	39	88.90585(2)	
zinc	Zn	30	65.409(4)	
zirconium	Zr	40	91.224(2)	g
*Element has no stable nuclides. However, three such elements (Th, Pa, and U) have a characteristic terrestrial isotopic composition, and for these an atomic mass is tabulated.				
†Commercially available Li materials have atomic weights that range between 6.939 and 6.996; if a more accurate value is required, it must be determined for the specific material.				
g Geological specimens are known in which the element has an isotopic composition outside the limits for normal material. The difference between the atomic mass of the element in such specimens and that given in the table may exceed the stated uncertainty.				
m Modified isotopic compositions may be found in commercially available material because it has been subjected to an undisclosed or inadvertent isotopic fractionation. Substantial deviations in the atomic mass of the element from that given in the table can occur.				
r Range in isotopic composition of normal terrestrial material prevents a more precise Ar(E) being given; the tabulated Ar(E) value and uncertainty should be applicable to normal material.				

Source: Adapted from *Pure and Applied Chemistry* 78, no. 11 (2005): 2051–66. © IUPAC (International Union of Pure and Applied Chemistry).

Selected Acid Dissociation Constants at 25°C

Name	Formula	K a1	pKa1	K a2	pKa2
Acetic acid	CH ₃ CO ₂ H	1.75×10^{-5}	4.756		
Arsenic acid	H ₃ AsO ₄	5.5×10^{-3}	2.26	1.7×10^{-7}	6.76
Benzoic acid	C ₆ H ₅ CO ₂ H	6.25×10^{-5}	4.204		
Boric acid	H ₃ BO ₃	$5.4 \times 10^{-10*}$	9.27*	$>1 \times 10^{-14*}$	>14*
Bromoacetic acid	CH ₂ BrCO ₂ H	1.3×10^{-3}	2.90		
Carbonic acid	H ₂ CO ₃	4.5×10^{-7}	6.35	4.7×10^{-11}	10.33
Chloroacetic acid	CH ₂ ClCO ₂ H	1.3×10^{-3}	2.87		
Chlorous acid	HClO ₂	1.1×10^{-2}	1.94		
Chromic acid	H ₂ CrO ₄	1.8×10^{-1}	0.74	3.2×10^{-7}	6.49
Citric acid	C ₆ H ₈ O ₇	7.4×10^{-4}	3.13	1.7×10^{-5}	4.76
Cyanic acid	HCNO	3.5×10^{-4}	3.46		
Dichloroacetic acid	CHCl ₂ CO ₂ H	4.5×10^{-2}	1.35		
Fluoroacetic acid	CH ₂ FCO ₂ H	2.6×10^{-3}	2.59		
Formic acid	CH ₂ O ₂	1.8×10^{-4}	3.75		
Hydrazoic acid	HN ₃	2.5×10^{-5}	4.6		
Hydrocyanic acid	HCN	6.2×10^{-10}	9.21		
Hydrofluoric acid	HF	6.3×10^{-4}	3.20		
Hydrogen selenide	H ₂ Se	1.3×10^{-4}	3.89	1.0×10^{-11}	11.0
Hydrogen sulfide	H ₂ S	8.9×10^{-8}	7.05	1×10^{-19}	19
Hydrogen telluride	H ₂ Te	$2.5 \times 10^{-3\ddagger}$	2.6^{\ddagger}	1×10^{-11}	11
Hypobromous acid	HBrO	2.8×10^{-9}	8.55		
Hypochlorous acid	HClO	4.0×10^{-8}	7.40		
Hypoiodous acid	HIO	3.2×10^{-11}	10.5		
Iodic acid	HIO ₃	1.7×10^{-1}	0.78		
Iodoacetic acid	CH ₂ ICO ₂ H	6.6×10^{-4}	3.18		

Name	Formula	K a1	pKa1	K a2	pKa2
Nitrous acid	HNO ₂	5.6×10^{-4}	3.25		
Oxalic acid	C ₂ H ₂ O ₄	5.6×10^{-2}	1.25	1.5×10^{-4}	3.81
Periodic acid	HIO ₄	2.3×10^{-2}	1.64		
Phenol	C ₆ H ₅ OH	1.0×10^{-10}	9.99		
Phosphoric acid	H ₃ PO ₄	6.9×10^{-3}	2.16	6.2×10^{-8}	7.21
Phosphorous acid	H ₃ PO ₃	$5.0 \times 10^{-2*}$	1.3*	$2.0 \times 10^{-7*}$	6.70*
Pyrophosphoric acid	H ₄ P ₂ O ₇	1.2×10^{-1}	0.91	7.9×10^{-3}	2.10
Resorcinol	C ₆ H ₄ (OH) ₂	4.8×10^{-10}	9.32	7.9×10^{-12}	11.1
Selenic acid	H ₂ SeO ₄	Strong	Strong	2.0×10^{-2}	1.7
Selenious acid	H ₂ SeO ₃	2.4×10^{-3}	2.62	4.8×10^{-9}	8.32
Sulfuric acid	H ₂ SO ₄	Strong	Strong	1.0×10^{-2}	1.99
Sulfurous acid	H ₂ SO ₃	1.4×10^{-2}	1.85	6.3×10^{-8}	7.2
meso-Tartaric acid	C ₄ H ₆ O ₆	6.8×10^{-4}	3.17	1.2×10^{-5}	4.91
Telluric acid	H ₂ TeO ₄	$2.1 \times 10^{-8\ddagger}$	$7.68\ddagger$	$1.0 \times 10^{-11\ddagger}$	$11.0\ddagger$
Tellurous acid	H ₂ TeO ₃	5.4×10^{-7}	6.27	3.7×10^{-9}	8.43
Trichloroacetic acid	CCl ₃ CO ₂ H	2.2×10^{-1}	0.66		
Trifluoroacetic acid	CF ₃ CO ₂ H	3.0×10^{-1}	0.52		

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Original source of data: *CRC Handbook of Chemistry and Physics*, 84th Edition (2004).

Solubility Constants for Compounds at 25°C.

Compound Name	Compound Formula	K _{sp}
Aluminum phosphate	AlPO ₄	9.84×10^{-21}
Barium bromate	Ba(BrO ₃) ₂	2.43×10^{-4}
Barium carbonate	BaCO ₃	2.58×10^{-9}
Barium chromate	BaCrO ₄	1.17×10^{-10}
Barium fluoride	BaF ₂	1.84×10^{-7}
Barium iodate	Ba(IO ₃) ₂	4.01×10^{-9}
Barium nitrate	Ba(NO ₃) ₂	4.64×10^{-3}
Barium sulphate	BaSO ₄	1.08×10^{-10}
Barium sulphite	BaSO ₃	5.0×10^{-10}
Beryllium hydroxide	Be(OH) ₂	6.92×10^{-22}
Bismuth arsenate	BiAsO ₄	4.43×10^{-10}
Bismuth iodide	BiI ₃	7.71×10^{-19}
Cadmium carbonate	CdCO ₃	1.0×10^{-12}
Cadmium fluoride	CdF ₂	6.44×10^{-3}
Cadmium hydroxide	Cd(OH) ₂	7.2×10^{-15}
Cadmium iodate	Cd(IO ₃) ₂	2.5×10^{-8}
Cadmium phosphate	Cd ₃ (PO ₄) ₂	2.53×10^{-33}
Cadmium sulphide	CdS	8.0×10^{-27}
Calcium carbonate	CaCO ₃	3.36×10^{-9}
Calcium fluoride	CaF ₂	3.45×10^{-11}
Calcium hydroxide	Ca(OH) ₂	5.02×10^{-6}
Calcium iodate	Ca(IO ₃) ₂	6.47×10^{-6}
Calcium phosphate	Ca ₃ (PO ₄) ₂	2.07×10^{-33}
Calcium sulphate	CaSO ₄	4.93×10^{-5}
Cesium perchlorate	CsClO ₄	3.95×10^{-3}
Cesium periodate	CsIO ₄	5.16×10^{-6}
Cobalt(II) arsenate	Co ₃ (AsO ₄) ₂	6.80×10^{-29}

Compound Name	Compound Formula	K _{sp}
Cobalt(II) hydroxide	Co(OH) ₂	5.92×10^{-15}
Cobalt(II) phosphate	Co ₃ (PO ₄) ₂	2.05×10^{-35}
Copper(I) bromide	CuBr	6.27×10^{-9}
Copper(I) chloride	CuCl	1.72×10^{-7}
Copper(I) cyanide	CuCN	3.47×10^{-20}
Copper(I) iodide	CuI	1.27×10^{-12}
Copper(I) thiocyanate	CuSCN	1.77×10^{-13}
Copper(II) arsenate	Cu ₃ (AsO ₄) ₂	7.95×10^{-36}
Copper(II) oxalate	CuC ₂ O ₄	4.43×10^{-10}
Copper(II) phosphate	Cu ₃ (PO ₄) ₂	1.40×10^{-37}
Copper(II) sulphide	CuS	6.3×10^{-36}
Europium(III) hydroxide	Eu(OH) ₃	9.38×10^{-27}
Gallium(III) hydroxide	Ga(OH) ₃	7.28×10^{-36}
Iron(II) carbonate	FeCO ₃	3.13×10^{-11}
Iron(II) fluoride	FeF ₂	2.36×10^{-6}
Iron(II) hydroxide	Fe(OH) ₂	4.87×10^{-17}
Iron(III) hydroxide	Fe(OH) ₃	2.79×10^{-39}
Iron(III) sulphide	FeS	6.3×10^{-18}
Lanthanum iodate	La(IO ₃) ₃	7.50×10^{-12}
Lead(II) bromide	PbBr ₂	6.60×10^{-6}
Lead(II) carbonate	PbCO ₃	7.40×10^{-14}
Lead(II) chloride	PbCl ₂	1.70×10^{-5}
Lead(II) fluoride	PbF ₂	3.3×10^{-8}
Lead(II) hydroxide	Pb(OH) ₂	1.43×10^{-20}
Lead(II) iodate	Pb(IO ₃) ₂	3.69×10^{-13}
Lead(II) iodide	PbI ₂	9.8×10^{-9}
Lead(II)selenite	PbSeO ₄	1.37×10^{-7}
Lead(II) sulphate	PbSO ₄	2.53×10^{-8}

Compound Name	Compound Formula	K _{sp}
Lead(II) sulphide	PbS	8.0×10^{-28}
Lithium carbonate	Li ₂ CO ₃	8.15×10^{-4}
Lithium fluoride	LiF	1.84×10^{-3}
Lithium phosphate	Li ₃ PO ₄	2.37×10^{-11}
Magnesium carbonate	MgCO ₃	6.82×10^{-6}
Magnesium fluoride	MgF ₂	5.16×10^{-11}
Magnesium hydroxide	Mg(OH) ₂	5.61×10^{-12}
Magnesium phosphate	Mg ₃ (PO ₄) ₂	1.04×10^{-24}
Manganese(II) carbonate	MnCO ₃	2.24×10^{-11}
Manganese(II) iodate	Mn(IO ₃) ₂	4.37×10^{-7}
Mercury(I) bromide	Hg ₂ Br ₂	6.40×10^{-23}
Mercury(I) carbonate	Hg ₂ CO ₃	3.6×10^{-17}
Mercury(I) chloride	Hg ₂ Cl ₂	1.43×10^{-18}
Mercury(I) fluoride	Hg ₂ F ₂	3.10×10^{-6}
Mercury(I) iodide	Hg ₂ I ₂	5.2×10^{-29}
Mercury(I) oxalate	Hg ₂ C ₂ O ₄	1.75×10^{-13}
Mercury(I) sulphate	Hg ₂ SO ₄	6.5×10^{-7}
Mercury(I) thiocyanate	Hg ₂ (SCN) ₂	3.2×10^{-20}
Mercury(II) bromide	HgBr ₂	6.2×10^{-20}
Mercury (II) iodide	HgI ₂	2.9×10^{-29}
Mercury(II) sulphide (red)	HgS	4×10^{-53}
Mercury(II) sulphide (black)	HgS	1.6×10^{-52}
Neodymium carbonate	Nd ₂ (CO ₃) ₃	1.08×10^{-33}
Nickel(II) carbonate	NiCO ₃	1.42×10^{-7}
Nickel(II) hydroxide	Ni(OH) ₂	5.48×10^{-16}
Nickel(II) iodate	Ni(IO ₃) ₂	4.71×10^{-5}
Nickel(II) phosphate	Ni ₃ (PO ₄) ₂	4.74×10^{-32}

Compound Name	Compound Formula	K _{sp}
Palladium(II) thiocyanate	Pd(SCN) ₂	4.39×10^{-23}
Potassium hexachloroplatinate	K ₂ PtCl ₆	7.48×10^{-6}
Potassium perchlorate	KClO ₄	1.05×10^{-2}
Potassium periodate	KIO ₄	3.71×10^{-4}
Praseodymium hydroxide	Pr(OH) ₃	3.39×10^{-24}
Rubidium perchlorate	RbClO ₄	3.00×10^{-3}
Scandium fluoride	ScF ₃	5.81×10^{-24}
Scandium hydroxide	Sc(OH) ₃	2.22×10^{-31}
Silver(I) acetate	AgCH ₃ CO ₂	1.94×10^{-3}
Silver(I) arsenate	Ag ₃ AsO ₄	1.03×10^{-22}
Silver(I) bromate	AgBrO ₃	5.38×10^{-5}
Silver(I) bromide	AgBr	5.35×10^{-13}
Silver(I) carbonate	Ag ₂ CO ₃	8.46×10^{-12}
Silver(I) chloride	AgCl	1.77×10^{-10}
Silver(I) chromate	Ag ₂ CrO ₄	1.12×10^{-12}
Silver(I) cyanide	AgCN	5.97×10^{-17}
Silver(I) iodate	AgIO ₃	3.17×10^{-8}
Silver(I) iodide	AgI	8.52×10^{-17}
Silver(I) oxalate	Ag ₂ C ₂ O ₄	5.40×10^{-12}
Silver(I) phosphate	Ag ₃ PO ₄	8.89×10^{-17}
Silver(I) sulphate	Ag ₂ SO ₄	1.20×10^{-5}
Silver(I) sulphide	Ag ₂ S	6.3×10^{-50}
Silver(I) sulfite	Ag ₂ SO ₃	1.50×10^{-14}
Silver(I) thiocyanate	AgSCN	1.03×10^{-12}
Strontium arsenate	Sr ₃ (AsO ₄) ₂	4.29×10^{-19}
Strontium carbonate	SrCO ₃	5.60×10^{-10}
Strontium fluoride	SrF ₂	4.33×10^{-9}

Compound Name	Compound Formula	K _{sp}
Strontium iodate	Sr(IO ₃) ₂	1.14×10^{-7}
Strontium sulphate	SrSO ₄	3.44×10^{-7}
Thallium(I) bromate	TlBrO ₃	1.10×10^{-4}
Thallium(I) bromide	TlBr	3.71×10^{-6}
Thallium(I) chloride	TlCl	1.86×10^{-4}
Thallium(I) chromate	Tl ₂ CrO ₄	8.67×10^{-13}
Thallium(I) iodate	TlIO ₃	3.12×10^{-6}
Thallium(I) iodide	TlI	5.54×10^{-8}
Thallium(I) thiocyanate	TlSCN	1.57×10^{-4}
Thallium(III) hydroxide	Tl(OH) ₃	1.68×10^{-44}
Tin(II) hydroxide	Sn(OH) ₂	5.45×10^{-27}
Tin(II) sulphide	SnS	1.0×10^{-25}
Yttrium carbonate	Y ₂ (CO ₃) ₃	1.03×10^{-31}
Yttrium fluoride	YF ₃	8.62×10^{-21}
Yttrium hydroxide	Y(OH) ₃	1.00×10^{-22}
Yttrium iodate	Y(IO ₃) ₃	1.12×10^{-10}
Zinc arsenate	Zn ₃ (AsO ₄) ₂	2.8×10^{-28}
Zinc carbonate	ZnCO ₃	1.46×10^{-10}
Zinc fluoride	ZnF ₂	3.04×10^{-2}
Zinc hydroxide	Zn(OH) ₂	3×10^{-17}
Zinc selenide	ZnSe	3.6×10^{-26}
Zinc sulphide (wurtzite)	ZnS	1.6×10^{-24}
Zinc sulphide (sphalerite)	ZnS	2.5×10^{-22}

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Original source of data: *CRC Handbook of Chemistry and*

Physics, 84th Edition (2004); sulfide data from *Lange's Handbook of Chemistry*, 15th Edition (1999).

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Aluminum:			
Al(s)	0.0	0.0	28.3
Al(g)	330.0	289.4	164.6
AlCl ₃ (s)	-704.2	-628.8	109.3
Al ₂ O ₃ (s)	-1675.7	-1582.3	50.9
Barium:			
Ba(s)	0.0	0.0	62.5
Ba(g)	180.0	146.0	170.2
BaO(s)	-548.0	-520.3	72.1
BaCO ₃ (s)	-1213.0	-1134.4	112.1
BaSO ₄ (s)	-1473.2	-1362.2	132.2
Beryllium:			
Be(s)	0.0	0.0	9.5
Be(g)	324.0	286.6	136.3
Be(OH) ₂ (s)	-902.5	-815.0	45.5
BeO(s)	-609.4	-580.1	13.8
Bismuth:			
Bi(s)	0.0	0.0	56.7
Bi(g)	207.1	168.2	187.0
Bromine:			

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Br(g)	111.9	82.4	175.0
Br₂(l)	0.0	0.0	152.2
Br⁻(aq)	-121.6	-104.0	82.4
Br₂(g)	30.9	3.1	245.5
HBr(g)	-36.3	-53.4	198.7
HBr(aq)	-121.6	-104.0	82.4
Cadmium:			
Cd(s)	0.0	0.0	51.8
Cd(g)	111.8	—	167.7
CdCl₂(s)	-391.5	-343.9	115.3
CdS(s)	-161.9	-156.5	64.9
Calcium:			
Ca(s)	0.0	0.0	41.6
Ca(g)	177.8	144.0	154.9
CaCl₂(s)	-795.4	-748.8	108.4
CaF₂(s)	-1228.0	-1175.6	68.5
Ca(OH)₂(s)	-985.2	-897.5	83.4
CaO(s)	-634.9	-603.3	38.1
CaSO₄(s)	-1434.5	-1322.0	106.5

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
CaCO₃(s, calcite)	-1207.6	-1129.1	91.7
CaCO₃(s, aragonite)	-1207.8	-1128.2	88.0
Carbon:			
C(s, graphite)	0.0	0.0	5.7
C(s, diamond)	1.9	2.9	2.4
C(s, fullerene—C₆₀)	2327.0	2302.0	426.0
C(s, fullerene—C₇₀)	2555.0	2537.0	464.0
C(g)	716.7	671.3	158.1
C(g, fullerene—C₆₀)	2502.0	2442.0	544.0
C(g, fullerene—C₇₀)	2755.0	2692.0	614.0
CBr₄(s)	29.4	47.7	212.5
CBr₄(g)	83.9	67.0	358.1
CCl₂F₂(g)	-477.4	-439.4	300.8
CCl₂O(g)	-219.1	-204.9	283.5
CCl₄(l)	-128.2	-62.6	216.2
CCl₄(g)	-95.7	-53.6	309.9

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
$\text{CF}_4(\text{g})$	-933.6	-888.3	261.6
$\text{CHCl}_3(\text{l})$	-134.1	-73.7	201.7
$\text{CHCl}_3(\text{g})$	-102.7	6.0	295.7
$\text{CH}_2\text{Cl}_2(\text{l})$	-124.2	—	177.8
$\text{CH}_2\text{Cl}_2(\text{g})$	-95.4	-68.9	270.2
$\text{CH}_3\text{Cl}(\text{g})$	-81.9	-58.5	234.6
$\text{CH}_4(\text{g})$	-74.6	-50.5	186.3
$\text{CH}_3\text{COOH}(\text{l})$	-484.3	-389.9	159.8
$\text{CH}_3\text{OH}(\text{l})$	-239.2	-166.6	126.8
$\text{CH}_3\text{OH}(\text{g})$	-201.0	-162.3	239.9
$\text{CH}_3\text{NH}_2(\text{l})$	-47.3	35.7	150.2
$\text{CH}_3\text{NH}_2(\text{g})$	-22.5	32.7	242.9
$\text{CH}_3\text{CN}(\text{l})$	40.6	86.5	149.6
$\text{CH}_3\text{CN}(\text{g})$	74.0	91.9	243.4
$\text{CO}(\text{g})$	-110.5	-137.2	197.7
$\text{CO}_2(\text{g})$	-393.5	-394.4	213.8
$\text{CS}_2(\text{l})$	89.0	64.6	151.3
$\text{CS}_2(\text{g})$	116.7	67.1	237.8

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
$C_2H_2(g)$	227.4	209.9	200.9
$C_2H_4(g)$	52.4	68.4	219.3
$C_2H_6(g)$	-84.0	-32.0	229.2
$C_3H_8(g)$	-103.8	-23.4	270.3
$C_3H_6O_3(s)$ (lactic acid)	-694.1	-522.9	142.3
$C_6H_6(l)$	49.1	124.5	173.4
$C_6H_6(g)$	82.9	129.7	269.2
$C_6H_{12}O_6(s)$ (glucose)	-1273.3	-910.4	212.1
$C_2H_5OH(l)$	-277.6	-174.8	160.7
$C_2H_5OH(g)$	-234.8	-167.9	281.6
$(CH_3)_2O(l)$	-203.3	—	—
$(CH_3)_2O(g)$	-184.1	-112.6	266.4
$CH_3CO_2^-(aq)$	-486.0	-369.3	86.6
$n\text{-}C_{12}H_{26}(l)$ (dodecane)	-350.9	28.1	490.6
Cesium:			
$Cs(s)$	0.0	0.0	85.2
$Cs(g)$	76.5	49.6	175.6

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
CsCl(s)	-443.0	-414.5	101.2
Chlorine:			
Cl(g)	121.3	105.3	165.2
Cl₂(g)	0.0	0.0	223.1
Cl⁻(aq)	-167.2	-131.2	56.5
HCl(g)	-92.3	-95.3	186.9
HCl(aq)	-167.2	-131.2	56.5
ClF₃(g)	-163.2	-123.0	281.6
Chromium:			
Cr(s)	0.0	0.0	23.8
Cr(g)	396.6	351.8	174.5
CrCl₃(s)	-556.5	-486.1	123.0
CrO₃(g)	-292.9	—	266.2
Cr₂O₃(s)	-1139.7	-1058.1	81.2
Cobalt:			
Co(s)	0.0	0.0	30.0
Co(g)	424.7	380.3	179.5
CoCl₂(s)	-312.5	-269.8	109.2
Copper:			
Cu(s)	0.0	0.0	33.2

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Cu(g)	337.4	297.7	166.4
CuCl(s)	-137.2	-119.9	86.2
CuCl₂(s)	-220.1	-175.7	108.1
CuO(s)	-157.3	-129.7	42.6
Cu₂O(s)	-168.6	-146.0	93.1
CuS(s)	-53.1	-53.6	66.5
Cu₂S(s)	-79.5	-86.2	120.9
CuCN(s)	96.2	111.3	84.5
Fluorine:			
F(g)	79.4	62.3	158.8
F⁻(aq)	-332.6	-278.8	-13.8
F₂(g)	0.0	0.0	202.8
HF(g)	-273.3	-275.4	173.8
HF(aq)	-332.6	-278.8	-13.8
Hydrogen:			
H(g)	218.0	203.3	114.7
H₂(g)	0.0	0.0	130.7
H⁺(aq)	0.0	0.0	0.0
Iodine:			

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
I(g)	106.8	70.2	180.8
I ⁻ (aq)	-55.2	-51.6	111.3
I ₂ (s)	0.0	0.0	116.1
I ₂ (g)	62.4	19.3	260.7
HI(g)	26.5	1.7	206.6
HI(aq)	-55.2	-51.6	111.3
Iron:			
Fe(s)	0.0	0.0	27.3
Fe(g)	416.3	370.7	180.5
Fe ²⁺ (aq)	-89.1	-78.9	-137.7
Fe ³⁺ (aq)	-48.5	-4.7	-315.9
FeCl ₂ (s)	-341.8	-302.3	118.0
FeCl ₃ (s)	-399.5	-334.0	142.3
FeO(s)	-272.0	-251.4	60.7
Fe ₂ O ₃ (s)	-824.2	-742.2	87.4
Fe ₃ O ₄ (s)	-1118.4	-1015.4	146.4
FeS ₂ (s)	-178.2	-166.9	52.9
FeCO ₃ (s)	-740.6	-666.7	92.9
Lead:			

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Pb(s)	0.0	0.0	64.8
Pb(g)	195.2	162.2	175.4
PbO(s, red or litharge)	-219.0	-188.9	66.5
PbO(s, yellow or massicot)	-217.3	-187.9	68.7
PbO ₂ (s)	-277.4	-217.3	68.6
PbCl ₂ (s)	-359.4	-314.1	136.0
PbS(s)	-100.4	-98.7	91.2
PbSO ₄ (s)	-920.0	-813.0	148.5
PbCO ₃ (s)	-699.1	-625.5	131.0
Pb(NO ₃) ₂ (s)	-451.9	—	—
Pb(NO ₃) ₂ (aq)	-416.3	-246.9	303.3
Lithium:			
Li(s)	0.0	0.0	29.1
Li(g)	159.3	126.6	138.8
Li ⁺ (aq)	-278.5	-293.3	13.4
LiCl(s)	-408.6	-384.4	59.3
Li ₂ O(s)	-597.9	-561.2	37.6
Magnesium:			

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Mg(s)	0.0	0.0	32.7
Mg(g)	147.1	112.5	148.6
MgCl ₂ (s)	-641.3	-591.8	89.6
MgO(s)	-601.6	-569.3	27.0
Mg(OH) ₂ (s)	-924.5	-833.5	63.2
MgSO ₄ (s)	-1284.9	-1170.6	91.6
MgS(s)	-346.0	-341.8	50.3
Manganese:			
Mn(s)	0.0	0.0	32.0
Mn(g)	280.7	238.5	173.7
MnCl ₂ (s)	-481.3	-440.5	118.2
MnO(s)	-385.2	-362.9	59.7
MnO ₂ (s)	-520.0	-465.1	53.1
KMnO ₄ (s)	-837.2	-737.6	171.7
MnO ₄ ⁻ (aq)	-541.4	-447.2	191.2
Mercury:			
Hg(l)	0.0	0.0	75.9
Hg(g)	61.4	31.8	175.0
HgCl ₂ (s)	-224.3	-178.6	146.0

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Hg₂Cl₂(s)	-265.4	-210.7	191.6
HgO(s)	-90.8	-58.5	70.3
HgS(s, red)	-58.2	-50.6	82.4
Hg₂(g)	108.8	68.2	288.1
Molybdenum:			
Mo(s)	0.0	0.0	28.7
Mo(g)	658.1	612.5	182.0
MoO₂(s)	-588.9	-533.0	46.3
MoO₃(s)	-745.1	-668.0	77.7
Nickel:			
Ni(s)	0.0	0.0	29.9
Ni(g)	429.7	384.5	182.2
NiCl₂(s)	-305.3	-259.0	97.7
Ni(OH)₂(s)	-529.7	-447.2	88.0
Nitrogen:			
N(g)	472.7	455.5	153.3
N₂(g)	0.0	0.0	191.6
NH₃(g)	-45.9	-16.4	192.8
NH₄⁺(aq)	-132.5	-79.3	113.4

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
$\text{N}_2\text{H}_4(\text{l})$	50.6	149.3	121.2
$\text{N}_2\text{H}_4(\text{g})$	95.4	159.4	238.5
$\text{NH}_4\text{Cl}(\text{s})$	-314.4	-202.9	94.6
$\text{NH}_4\text{OH}(\text{l})$	-361.2	-254.0	165.6
$\text{NH}_4\text{NO}_3(\text{s})$	-365.6	-183.9	151.1
$(\text{NH}_4)_2\text{SO}_4(\text{s})$	-1180.9	-901.7	220.1
$\text{NO}(\text{g})$	91.3	87.6	210.8
$\text{NO}_2(\text{g})$	33.2	51.3	240.1
$\text{N}_2\text{O}(\text{g})$	81.6	103.7	220.0
$\text{N}_2\text{O}_4(\text{l})$	-19.5	97.5	209.2
$\text{N}_2\text{O}_4(\text{g})$	11.1	99.8	304.4
$\text{HNO}_2(\text{g})$	-79.5	-46.0	254.1
$\text{HNO}_3(\text{l})$	-174.1	-80.7	155.6
$\text{HNO}_3(\text{g})$	-133.9	-73.5	266.9
$\text{HNO}_3(\text{aq})$	-207.4	-111.3	146.4
$\text{NF}_3(\text{g})$	-132.1	-90.6	260.8
$\text{HCN}(\text{l})$	108.9	125.0	112.8
$\text{HCN}(\text{g})$	135.1	124.7	201.8

Osmium:

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Os(s)	0.0	0.0	32.6
Os(g)	791.0	745.0	192.6
OsO ₄ (s)	-394.1	-304.9	143.9
OsO ₄ (g)	-337.2	-292.8	293.8
Oxygen:			
O(g)	249.2	231.7	161.1
O ₂ (g)	0.0	0.0	205.2
O ₃ (g)	142.7	163.2	238.9
OH ⁻ (aq)	-230.0	-157.2	-10.8
H ₂ O(l)	-285.8	-237.1	70.0
H ₂ O(g)	-241.8	-228.6	188.8
H ₂ O ₂ (l)	-187.8	-120.4	109.6
H ₂ O ₂ (g)	-136.3	-105.6	232.7
Phosphorus:			
P(s, white)	0.0	0.0	41.1
P(s, red)	-17.6	-12.5	22.8
P(s, black)	-39.3	—	—
P(g, white)	316.5	280.1	163.2
P ₂ (g)	144.0	103.5	218.1

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
P₄(g)	58.9	24.4	280.0
PCl₃(l)	-319.7	-272.3	217.1
PCl₃(g)	-287.0	-267.8	311.8
POCl₃(l)	-597.1	-520.8	222.5
POCl₃(g)	-558.5	-512.9	325.5
PCl₅(g)	-374.9	-305.0	364.6
PH₃(g)	5.4	13.5	210.2
H₃PO₄(s)	-1284.4	-1124.3	110.5
H₃PO₄(l)	-1271.7	-1123.6	150.8
Potassium:			
K(s)	0.0	0.0	64.7
K(g)	89.0	60.5	160.3
KBr(s)	-393.8	-380.7	95.9
KCl(s)	-436.5	-408.5	82.6
KClO₃(s)	-397.7	-296.3	143.1
K₂O(s)	-361.5	-322.1	94.1
K₂O₂(s)	-494.1	-425.1	102.1
KNO₂(s)	-369.8	-306.6	152.1
KNO₃(s)	-494.6	-394.9	133.1

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
KSCN(s)	-200.2	-178.3	124.3
K ₂ CO ₃ (s)	-1151.0	-1063.5	155.5
K ₂ SO ₄ (s)	-1437.8	-1321.4	175.6
Rubidium:			
Rb(s)	0.0	0.0	76.8
Rb(g)	80.9	53.1	170.1
RbCl(s)	-435.4	-407.8	95.9
Selenium:			
Se(s, gray)	0.0	0.0	42.4
Se(g, gray)	227.1	187.0	176.7
H ₂ Se(g)	29.7	15.9	219.0
Silicon:			
Si(s)	0.0	0.0	18.8
Si(g)	450.0	405.5	168.0
SiCl ₄ (l)	-687.0	-619.8	239.7
SiCl ₄ (g)	-657.0	-617.0	330.7
SiH ₄ (g)	34.3	56.9	204.6
SiC(s, cubic)	-65.3	-62.8	16.6
SiC(s, hexagonal)	-62.8	-60.2	16.5

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Silver:			
Ag(s)	0.0	0.0	42.6
Ag(g)	284.9	246.0	173.0
Ag⁺(aq)	105.6	77.1	72.7
AgBr(s)	-100.4	-96.9	107.1
AgCl(s)	-127.0	-109.8	96.3
AgNO₃(s)	-124.4	-33.4	140.9
Ag₂O(s)	-31.1	-11.2	121.3
Ag₂S(s)	-32.6	-40.7	144.0
Sodium:			
Na(s)	0.0	0.0	51.3
Na(g)	107.5	77.0	153.7
Na⁺(aq)	-240.1	-261.9	59.0
NaF(s)	-576.6	-546.3	51.1
NaF(aq)	-572.8	-540.7	45.2
NaCl(s)	-411.2	-384.1	72.1
NaCl(aq)	-407.3	-393.1	115.5
NaBr(s)	-361.1	-349.0	86.8
NaBr(g)	-143.1	-177.1	241.2

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
NaBr(aq)	-361.7	-365.8	141.4
NaO₂(s)	-260.2	-218.4	115.9
Na₂O(s)	-414.2	-375.5	75.1
Na₂O₂(s)	-510.9	-447.7	95.0
NaCN(s)	-87.5	-76.4	115.6
NaNO₃(aq)	-447.5	-373.2	205.4
NaNO₃(s)	-467.9	-367.0	116.5
NaN₃(s)	21.7	93.8	96.9
Na₂CO₃(s)	-1130.7	-1044.4	135.0
Na₂SO₄(s)	-1387.1	-1270.2	149.6
Sulphur:			
S(s, rhombic)	0.0	0.0	32.1
S(g, rhombic)	277.2	236.7	167.8
SO₂(g)	-296.8	-300.1	248.2
SO₃(g)	-395.7	-371.1	256.8
SO₄²⁻(aq)	-909.3	-744.5	20.1
SOCl₂(g)	-212.5	-198.3	309.8
H₂S(g)	-20.6	-33.4	205.8
H₂SO₄(aq)	-909.3	-744.5	20.1

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Tin:			
Sn(s, white)	0.0	0.0	51.2
Sn(s, gray)	-2.1	0.1	44.1
Sn(g, white)	301.2	266.2	168.5
SnCl ₄ (l)	-511.3	-440.1	258.6
SnCl ₄ (g)	-471.5	-432.2	365.8
SnO ₂ (s)	-557.6	-515.8	49.0
Titanium:			
Ti(s)	0.0	0.0	30.7
Ti(g)	473.0	428.4	180.3
TiCl ₂ (s)	-513.8	-464.4	87.4
TiCl ₃ (s)	-720.9	-653.5	139.7
TiCl ₄ (l)	-804.2	-737.2	252.3
TiCl ₄ (g)	-763.2	-726.3	353.2
TiO ₂ (s)	-944.0	-888.8	50.6
Uranium:			
U(s)	0.0	0.0	50.2
U(g)	533.0	488.4	199.8
UO ₂ (s)	-1085.0	-1031.8	77.0

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
UO₂(g)	-465.7	-471.5	274.6
UF₄(s)	-1914.2	-1823.3	151.7
UF₄(g)	-1598.7	-1572.7	368.0
UF₆(s)	-2197.0	-2068.5	227.6
UF₆(g)	-2147.4	-2063.7	377.9
Vanadium:			
V(s)	0.0	0.0	28.9
V(g)	514.2	754.4	182.3
VCl₃(s)	-580.7	-511.2	131.0
VCl₄(l)	-569.4	-503.7	255.0
VCl₄(g)	-525.5	-492.0	362.4
V₂O₅(s)	-1550.6	-1419.5	131.0
Zinc:			
Zn(s)	0.0	0.0	41.6
Zn(g)	130.4	94.8	161.0
ZnCl₂(s)	-415.1	-369.4	111.5
Zn(NO₃)₂(s)	-483.7	—	—
ZnS(s, sphalerite)	-206.0	-201.3	57.7
ZnSO₄(s)	-982.8	-871.5	110.5

Substance	ΔH_f° (kJ/mol)	ΔG_f° (kJ/mol)	S° (J/mol K)
Zirconium:			
Zr(s)	0.0	0.0	39.0
Zr(g)	608.8	566.5	181.4
ZrCl₂(s)	-502.0	-386	110
ZrCl₄(s)	-980.5	-889.9	181.6

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Original Source of data: CRC Handbook of Chemistry and Physics, 84th Edition (2004).

Standard Cathode (Reduction) Half-Reaction	Standard Reduction Potential E° (volts)
$\text{Li}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Li}(\text{s})$	-3.040
$\text{Ba}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ba}(\text{s})$	-2.92
$\text{Rb}^+ + \text{e}^- \rightleftharpoons \text{Rb}(\text{s})$	-2.98
$\text{K}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{K}(\text{s})$	-2.93
$\text{Cs}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cs}(\text{s})$	-2.92
$\text{Ba}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Ba}(\text{s})$	-2.91
$\text{Sr}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sr}(\text{s})$	-2.89
$\text{Ca}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Ca}(\text{s})$	-2.84
$\text{Na}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Na}(\text{s})$	-2.713
$\text{Mg}(\text{OH})_2(\text{s}) + 2\text{e}^- \rightleftharpoons \text{Mg}(\text{s}) + 2\text{OH}^-$	-2.687
$\text{La}^{3+} + 3\text{e}^- \rightleftharpoons \text{La}(\text{s})$	-2.38
$\text{Mg}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Mn}(\text{s})$	-2.356
$\text{Ce}^{3+} + 3\text{e}^- \rightleftharpoons \text{Ce}(\text{s})$	-2.336
$\text{Al}(\text{OH})_4^- + 3\text{e}^- \rightleftharpoons \text{Al}(\text{s}) + 4\text{OH}^-$	-2.310
$\text{AlF}_6^{3-} + 3\text{e}^- \rightleftharpoons \text{Al}(\text{s}) + 6\text{F}^-$	-2.07
$\text{Be}^{2+} + 2\text{e}^- \rightleftharpoons \text{Be}(\text{s})$	-1.99
$\text{B}(\text{OH})_4^- + 3\text{e}^- \rightleftharpoons \text{B}(\text{s}) + 4\text{OH}^-$	-1.811
$\text{U}^{3+} + 3\text{e}^- \rightleftharpoons \text{U}(\text{s})$	-1.66

$\text{Al}^{3+}(\text{aq}) + 3\text{e}^{-} \rightleftharpoons \text{Al}(\text{s})$	-1.676
$\text{SiF}_6^{2-} + 4\text{e}^{-} \rightleftharpoons \text{Si}(\text{s}) + 6\text{F}^{-}$	-1.37
$\text{Zn}(\text{CN})_4^{2-} + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s}) + 4\text{CN}^{-}$	-1.34
$\text{Zn}(\text{OH})_4^{2-} + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s}) + 4\text{OH}^{-}$	-1.285
$\text{Mn}^{2+} + 2\text{e}^{-} \rightleftharpoons \text{Mn}(\text{s})$	-1.17
$\text{V}^{2+} + 2\text{e}^{-} \rightleftharpoons \text{V}(\text{s})$	-1.13
$2\text{SO}_3^{2-} + 2\text{H}_2\text{O}(\text{l}) + 2\text{e}^{-} \rightleftharpoons \text{S}_2\text{O}_4^{2-} + 4\text{OH}^{-}$	-1.13
$\text{Zn}(\text{NH}_3)_4^{2+} + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s}) + 4\text{NH}_3$	-1.04
$\text{O}_2(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{O}_2^{-}(\text{aq})$	-1.0
$\text{Cd}(\text{CN})_4^{2-} + 2\text{e}^{-} \rightleftharpoons \text{Cd}(\text{s}) + 4\text{CN}^{-}$	-0.943
$\text{MoO}_4^{2-} + 4\text{H}_2\text{O}(\text{l}) + 6\text{e}^{-} \rightleftharpoons \text{Mo}(\text{s}) + 8\text{OH}^{-}$	-0.913
$\text{SiO}_2(\text{s}) + 4\text{H}^{+} + 4\text{e}^{-} \rightleftharpoons \text{Si}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	-0.909
$\text{SO}_4^{2-} + \text{H}_2\text{O}(\text{l}) + 2\text{e}^{-} \rightleftharpoons \text{SO}_3^{2-} + 2\text{OH}^{-}$	-0.936
$\text{Cr}^{2+} + 2\text{e}^{-} \rightleftharpoons \text{Cr}(\text{s})$	-0.90
$\text{B}(\text{OH})_3 + 3\text{H}^{+} + 3\text{e}^{-} \rightleftharpoons \text{B}(\text{s}) + 3\text{H}_2\text{O}(\text{l})$	-0.890
$2\text{H}_2\text{O}(\text{l}) + 2\text{e}^{-} \rightleftharpoons \text{H}_2(\text{g}) + 2\text{OH}^{-}(\text{aq})$	-0.828
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s})$	-0.7618
$\text{Co}(\text{OH})_2(\text{s}) + 2\text{e}^{-} \rightleftharpoons \text{Co}(\text{s}) + 2\text{OH}^{-}$	-0.746
$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-} \rightleftharpoons \text{Cr}(\text{s})$	-0.424

$\text{Ni}(\text{OH})_2 + 2e^- \rightleftharpoons \text{Ni}(\text{s}) + 2\text{OH}^-$	-0.72
$\text{Ag}_2\text{S}(\text{s}) + 2e^- \rightleftharpoons 2\text{Ag}(\text{s}) + \text{S}^{2-}$	-0.71
$\text{Se}(\text{s}) + 2e^- \rightleftharpoons \text{Se}^{2-}$	-0.67 in 1 M NaOH
$\text{Cd}(\text{NH}_3)_4^{2+} + 2e^- \rightleftharpoons \text{Cd}(\text{s}) + 4\text{NH}_3$	-0.622
$2\text{SO}_3^{2-} + 3\text{H}_2\text{O}(\text{l}) + 4e^- \rightleftharpoons \text{S}_2\text{O}_3^{2-} + 6\text{OH}^-$	-0.576 in 1 M NaOH
$\text{U}^{4+} + e^- \rightleftharpoons \text{U}^{3+}$	-0.52
$\text{SiO}_2(\text{s}) + 8\text{H}^+ + 8e^- \rightleftharpoons \text{SiH}_4(\text{g}) + 2\text{H}_2\text{O}(\text{l})$	-0.516
$\text{Sb} + 3\text{H}^+ + 3e^- \rightleftharpoons \text{SbH}_3(\text{g})$	-0.510
$\text{H}_3\text{PO}_3 + 2\text{H}^+ + 2e^- \rightleftharpoons \text{H}_3\text{PO}_2 + \text{H}_2\text{O}(\text{l})$	-0.50
$\text{Ni}(\text{NH}_3)_6^{2+} + 2e^- \rightleftharpoons \text{Ni}(\text{s}) + 6\text{NH}_3$	-0.49
$2\text{CO}_2(\text{g}) + 2\text{H}^+ + 2e^- \rightleftharpoons \text{H}_2\text{C}_2\text{O}_4$	-0.481
$\text{Cr}^{3+} + e^- \rightleftharpoons \text{Cr}^{2+}$	-0.424
$\text{Fe}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Fe}(\text{s})$	-0.44
$\text{S}(\text{s}) + 2e^- \rightleftharpoons \text{S}^{2-}$	-0.407
$\text{Cd}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Cd}(\text{s})$	-0.4030
$\text{Ag}(\text{NH}_3)_2^+ + e^- \rightleftharpoons \text{Ag}(\text{s}) + 2\text{NH}_3$	-0.373
$\text{Ti}^{3+} + e^- \rightleftharpoons \text{Ti}^{2+}$	-0.37
$\text{PbSO}_4(\text{s}) + 2e^- \rightleftharpoons \text{Pb}(\text{s}) + \text{SO}_4^{2-}$	-0.356
$\text{Co}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Co}(\text{s})$	-0.277

$2\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{S}_2\text{O}_6^{2-} + 2\text{H}_2\text{O}(\text{l})$	-0.25
$\text{N}_2(\text{g}) + 5\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{N}_2\text{H}_5^+$	-0.23
$\text{H}_3\text{PO}_4 + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_3\text{PO}_3 + \text{H}_2\text{O}(\text{l})$	-0.28
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Ni}(\text{s})$	-0.257
$\text{V}^{3+} + \text{e}^- \rightleftharpoons \text{V}^{2+}$	-0.255
$\text{As} + 3\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{AsH}_3(\text{g})$	-0.225
$\text{CO}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{HCO}_2\text{H}$	-0.20
$\text{Mo}^{3+} + 3\text{e}^- \rightleftharpoons \text{Mo}(\text{s})$	-0.2
$\text{Sn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	-0.19 in 1 M HCl
$\text{Ti}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ti}(\text{s})$	-0.163
$\text{MoO}_2(\text{s}) + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{Mo}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	-0.152
$\text{AgI}(\text{s}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + \text{I}^-$	-0.152
$\text{Sn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}(\text{s})$	-0.14
$\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Pb}(\text{s})$	-0.126
$\text{CrO}_4^{2-} + 4\text{H}_2\text{O}(\text{l}) + 3\text{e}^- \rightleftharpoons 2\text{Cr}(\text{OH})_4^- + 4\text{OH}^-$	-0.13 in 1 M NaOH
$\text{WO}_2(\text{s}) + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{W}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	-0.119
$\text{Se}(\text{s}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{Se}(\text{g})$	-0.115
$\text{CO}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CO}(\text{g}) + \text{H}_2\text{O}(\text{l})$	-0.106

$\text{WO}_3(\text{s}) + 6\text{H}^+ + 6\text{e}^- \rightleftharpoons \text{W}(\text{s}) + 3\text{H}_2\text{O}(\text{l})$	-0.090
$\text{Hg}_2\text{I}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{Hg}(\text{l}) + 2\text{I}^-$	-0.0405
$\text{Fe}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	-0.037
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	0.00
$\text{P}(\text{s,white}) + 3\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{PH}_3(\text{g})$	0.06
$\text{AgBr}(\text{s}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + \text{Br}^-$	0.071
$\text{S}_4\text{O}_6^{2-} + 2\text{e}^- \rightleftharpoons 2\text{S}_2\text{O}_3^{2-}$	0.080
$\text{Co}(\text{NH}_3)_6^{3+} + \text{e}^- \rightleftharpoons \text{Co}(\text{NH}_3)_6^{2+}$	0.1
$\text{Ru}(\text{NH}_3)_6^{3+} + \text{e}^- \rightleftharpoons \text{Ru}(\text{s}) + \text{Ru}(\text{NH}_3)_6^{2+}$	0.10
$\text{S}(\text{s}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{S}$	0.144
$\text{Sn}^{4+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}(\text{aq})$	0.154
$\text{Cu}^{2+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu}^+(\text{aq})$	0.159
$\text{UO}_2^{2+} + \text{e}^- \rightleftharpoons \text{UO}_2^+$	0.16
$\text{Co}(\text{OH})_3(\text{s}) + \text{e}^- \rightleftharpoons \text{Co}(\text{OH})_2(\text{s}) + \text{OH}^-$	0.17
$\text{ClO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightleftharpoons \text{ClO}_3^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.17
$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{SO}_3^{2-} + \text{H}_2\text{O}(\text{l})$	0.172
$\text{BiCl}_4^- + 3\text{e}^- \rightleftharpoons \text{Bi}(\text{s}) + 4\text{Cl}^-$	0.199
$\text{SbO}^+ + 2\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{Sb}(\text{s}) + \text{H}_2\text{O}(\text{l})$	0.212

$\text{AgCl(s)} + \text{e}^- \rightleftharpoons \text{Ag(s)} + \text{Cl}^-(\text{aq})$	0.2223
$\text{HCHO} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CH}_3\text{OH}$	0.2323
$\text{HAsO}_2 + 3\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{As(s)} + 2\text{H}_2\text{O(l)}$	0.240
$\text{Ru}^{3+} + \text{e}^- \rightleftharpoons \text{Ru}^{2+}$	0.249
$\text{IO}_3^- + 3\text{H}_2\text{O(l)} + 6\text{e}^- \rightleftharpoons \text{I}^- + 6\text{OH}^-$	0.257
$\text{Hg}_2\text{Cl}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{Hg(l)} + 2\text{Cl}^-$	0.2682
$\text{UO}_2^{+} + 4\text{H}^+ + \text{e}^- \rightleftharpoons \text{U}^{4+} + 2\text{H}_2\text{O(l)}$	0.27
$\text{Bi}^{3+} + 3\text{e}^- \rightleftharpoons \text{Bi(s)}$	0.317
$\text{UO}_2^{2+} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{U}^{4+} + 2\text{H}_2\text{O(l)}$	0.327
$\text{VO}^{2+} + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{V}^{3+} + \text{H}_2\text{O(l)}$	0.337
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu(s)}$	0.3419
$\text{ClO}_3^-(\text{aq}) + \text{H}_2\text{O(l)} + 2\text{e}^- \rightleftharpoons \text{ClO}_2^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.35
$\text{Fe}(\text{CN})_6^{3-} + \text{e}^- \rightleftharpoons \text{Fe}(\text{CN})_6^{4-}$	0.356
$\text{O}_2(\text{g}) + 2\text{H}_2\text{O(l)} + 4\text{e}^- \rightleftharpoons 4\text{OH}^-$	0.401
$\text{ClO}^- + \text{H}_2\text{O(l)} + \text{e}^- \rightleftharpoons \frac{1}{2}\text{Cl}_2(\text{g}) + 2\text{OH}^-$	0.421 in 1 M NaOH
$\text{Ag}_2\text{C}_2\text{O}_4(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{Ag(s)} + \text{C}_2\text{O}_4^{2-}$	0.47
$\text{Cu}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Cu(s)}$	0.52
$\text{I}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{I}^-(\text{aq})$	0.5355

$\text{I}_3^- + 2e^- \rightleftharpoons 3\text{I}^-$	0.536
$\text{Ga}^{3+} + 3e^- \rightleftharpoons \text{Ga(s)}$	-0.56
$\text{Cu}^{2+} + \text{Cl}^- + e^- \rightleftharpoons \text{CuCl(s)}$	0.559
$\text{S}_2\text{O}_6^{2-} + 4\text{H}^+ + 2e^- \rightleftharpoons 2\text{H}_2\text{SO}_3$	0.569
$\text{H}_3\text{AsO}_4 + 2\text{H}^+ + 2e^- \rightleftharpoons \text{HAsO}_2 + 2\text{H}_2\text{O(l)}$	0.560
$\text{ClO}_2^-(\text{aq}) + \text{H}_2\text{O(l)} + 2e^- \rightleftharpoons \text{ClO}^-(\text{aq}) + 2\text{OH}^-(\text{aq})$	0.59
$\text{MnO}_4^- + 2\text{H}_2\text{O(l)} + 3e^- \rightleftharpoons \text{MnO}_2(\text{s}) + 4\text{OH}^-$	0.60
$\text{Sb}_2\text{O}_5(\text{s}) + 6\text{H}^+ + 4e^- \rightleftharpoons 2\text{SbO}^+ + 3\text{H}_2\text{O(l)}$	0.605
$\text{PtCl}_6^{2-} + 2e^- \rightleftharpoons \text{PtCl}_4^{2-} + 2\text{Cl}^-$	0.68
$\text{RuO}_2(\text{s}) + 4\text{H}^+ + 4e^- \rightleftharpoons \text{Ru(s)} + 2\text{H}_2\text{O(l)}$	0.68
$\text{O}_2(\text{g}) + 2\text{H}^+ + 2e^- \rightleftharpoons \text{H}_2\text{O}_2$	0.695
$\text{PtCl}_4^{2-} + 2e^- \rightleftharpoons \text{Pt(s)} + 4\text{Cl}^-$	0.73
$\text{H}_2\text{SeO}_3 + 4\text{H}^+ + 4e^- \rightleftharpoons \text{Se(s)} + 3\text{H}_2\text{O(l)}$	0.74
$\text{Tl}^{3+} + 3e^- \rightleftharpoons \text{Tl(s)}$	0.742
$\text{Fe}^{3+}(\text{aq}) + e^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	0.771
$\text{Hg}_2^{2+}(\text{aq}) + 2e^- \rightleftharpoons 2\text{Hg(l)}$	0.7960
$\text{Ag}^+(\text{aq}) + e^- \rightleftharpoons \text{Ag(s)}$	0.7996
$\text{Hg}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Hg(l)}$	0.8535

$\text{Cu}^{2+} + \text{I}^- + e^- \rightleftharpoons \text{CuI(s)}$	0.86
$\text{Ru(CN)}_6^{3-} + e^- \rightleftharpoons \text{Ru(s)} + \text{Ru(CN)}_6^{4-}$	0.86
$\text{ClO}^- + \text{H}_2\text{O(l)} + 2e^- \rightleftharpoons \text{Cl}^- + 2\text{OH}^-$	0.890 in 1 M NaOH
$2\text{Hg}^{2+}(\text{aq}) + 2e^- \rightleftharpoons \text{Hg}_2^{2+}(\text{aq})$	0.911
$\text{HgO(s)} + 2\text{H}^+ + 2e^- \rightleftharpoons \text{Hg(l)} + \text{H}_2\text{O(l)}$	0.926
$\text{NO}_3^- + 3\text{H}^+ + 2e^- \rightleftharpoons \text{HNO}_2 + \text{H}_2\text{O(l)}$	0.94
$\text{MnO}_2(\text{s}) + 4\text{H}^+ + e^- \rightleftharpoons \text{Mn}^{3+}(\text{aq}) + \text{H}_2\text{O(l)}$	0.95
$\text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3e^- \rightleftharpoons \text{NO(g)} + 2\text{H}_2\text{O(l)}$	0.96
$\text{HIO} + \text{H}^+ + 2e^- \rightleftharpoons \text{I}^- + \text{H}_2\text{O(l)}$	0.985
$\text{HNO}_2 + \text{H}^+ + e^- \rightleftharpoons \text{NO(g)} + \text{H}_2\text{O(l)}$	0.996
$\text{VO}_2^{2+} + 2\text{H}^+ + e^- \rightleftharpoons \text{VO}^{2+} + \text{H}_2\text{O(l)}$	1.000
$\text{AuCl}_4^- + 3e^- \rightleftharpoons \text{Au(s)} + 4\text{Cl}^-$	1.002
$\text{NO}_2(\text{g}) + \text{H}^+(\text{aq}) + e^- \rightleftharpoons \text{HNO}_2(\text{aq})$	1.07
$\text{Br}_2(\text{l}) + 2e^- \rightleftharpoons 2\text{Br}^-(\text{aq})$	1.087
$\text{Fe(phen)}_6^{3+} + e^- \rightleftharpoons \text{Fe(phen)}_6^{2+}$	1.147
$\text{SeO}_4^{3-} + 4\text{H}^+ + e^- \rightleftharpoons \text{H}_2\text{SeO}_3 + \text{H}_2\text{O(l)}$	1.151
$\text{ClO}_3^- + 2\text{H}^+ + e^- \rightleftharpoons \text{ClO}_2(\text{g}) + \text{H}_2\text{O}$	1.175
$\text{ClO}_3^- + 3\text{H}^+ + 2e^- \rightleftharpoons \text{HClO}_2 + \text{H}_2\text{O}$	1.181

$\text{IO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightleftharpoons \frac{1}{2}\text{I}_2(\text{s}) + 3\text{H}_2\text{O}(\text{l})$	1.195
$\text{Pt}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pt}(\text{s})$	1.2
$\text{ClO}_4^- + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{ClO}_3^- + \text{H}_2\text{O}$	1.201
$\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$	1.229
$\text{MnO}_2(\text{s}) + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 2\text{H}_2\text{O}(\text{l})$	1.23
$\text{Tl}^{3+} + 2\text{e}^- \rightleftharpoons \text{Tl}^+$	0.77 in 1 M HCl
$2\text{HNO}_2 + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{N}_2\text{O}(\text{g}) + 3\text{H}_2\text{O}(\text{l})$	1.297
$\text{HOBr} + \text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Br}^- + \text{H}_2\text{O}(\text{l})$	1.341
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	1.36
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+} + 7\text{H}_2\text{O}(\text{l})$	1.36
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$	1.396
$\text{Au}^{3+} + 2\text{e}^- \rightleftharpoons \text{Au}^+$	1.36
$\text{Hg}_2\text{Br}_2(\text{s}) + 2\text{e}^- \rightleftharpoons 2\text{Hg}(\text{l}) + 2\text{Br}^-$	1.392
$\text{Ce}^{4+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ce}^{3+}(\text{aq})$	1.44
$\text{PbO}_2(\text{s}) + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$	1.46
$\text{BrO}_3^- + 6\text{H}^+ + 6\text{e}^- \rightleftharpoons \text{Br}^- + 3\text{H}_2\text{O}$	1.478
$\text{Mn}^{3+} + \text{e}^- \rightleftharpoons \text{Mn}^{2+}$	1.5
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	1.51

$\text{BrO}_3^- + 6\text{H}^+ + 5\text{e}^- \rightleftharpoons \frac{1}{2}\text{Br}_2(\text{l}) + 3\text{H}_2\text{O}$	1.5
$\text{Au}^{3+} + 3\text{e}^- \rightleftharpoons \text{Au}(\text{s})$	1.52
$2\text{NO}(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{N}_2\text{O}(\text{g}) + \text{H}_2\text{O}(\text{l})$	1.59
$\text{HOBr} + \text{H}^+ + \text{e}^- \rightleftharpoons \frac{1}{2}\text{Br}^- + \text{H}_2\text{O}(\text{l})$	1.604
$\text{HClO}_2 + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{HOCl} + \text{H}_2\text{O}$	1.64
$\text{PbO}_2(\text{s}) + 4\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	1.690
$\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	1.70
$\text{Ce}^{4+} + \text{e}^- \rightleftharpoons \text{Ce}^{3+}$	1.72
$\text{N}_2\text{O}(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{N}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$	1.77
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}(\text{l})$	1.763
$\text{Au}^+ + \text{e}^- \rightleftharpoons \text{Au}(\text{s})$	1.83
$\text{Co}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Co}^{2+}(\text{aq})$	1.92
$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightleftharpoons 2\text{SO}_4^{2-}$	1.96
$\text{O}_3(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$	2.07
$\text{BaO}(\text{s}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Ba}(\text{s}) + \text{H}_2\text{O}(\text{l})$	2.365
$\text{F}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{F}^-(\text{aq})$	2.87
$\text{F}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{HF}$	3.053

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Term	Definition	Section of Book
abbreviated electron configuration	An electron configuration that uses one of the noble gases to represent the core of electrons up to that element.	Organization of Electrons in Atoms
absolute zero	The minimum possible temperature, labeled 0 K (zero kelvins)	Other Units: Temperature and Density
acid	An ionic compound of the H^+ cation dissolved in water	Acids
acid	A compound that increases the amount of H^+ ions in an aqueous solution	Neutralization Reactions
acid dissociation constant (K_a)	The equilibrium constant for the dissociation of a weak acid into ions	Some Special Types of Equilibria
acid salt	An ionic compound whose aqueous solution is slightly acidic	Strong and Weak Acids and Bases and Their Salts
activated complex	See transition state	Reaction Mechanisms
activation energy (E_a)	The minimum amount of kinetic energy molecules must possess for an effective collision to occur	Factors that Affect the Rate of Reactions
active site	Area of enzymatic action where substrate molecules react	Catalysis
activity series	A list of elements that will replace elements below them in single-replacement reactions	Types of Chemical Reactions: Single- and Double-Displacement Reactions
actual yield	The amount that is actually produced in a chemical reaction	Yields
addition reaction	A reaction where atoms are added across a double or triple bond	Hydrocarbons

adhesion	The tendency of a substance to interact with other substances because of intermolecular forces	Properties of Liquids
adsorb	Bind to the surface of another substance	Catalysis
alcohol	An organic compound that contains an OH functional group	Alkyl halides and alcohols
aldehyde	A compound that has a carbonyl functional group at the end of a chain of C atoms	Other Oxygen-Containing Functional Groups
aliphatic hydrocarbons	A hydrocarbon based on chains of C atoms	Hydrocarbons
alkaline battery	A type of dry cell that contains an alkaline (i.e., basic) moist paste, rather than an acidic paste	Applications of Redox Reactions: Voltaic Cells
alkane	An aliphatic hydrocarbon with only single covalent bonds	Hydrocarbons
alkene	An aliphatic hydrocarbon that contains a C–C double bond	Hydrocarbons
alkyl halide	An organic compound that contains a halogen atom	Alkyl halides and alcohols
alkyne	An aliphatic hydrocarbon that contains a C–C triple bond	Hydrocarbons
alpha particle	A type of radioactive emission equivalent to a helium nucleus	Radioactivity
amide group	A functional group that is the combination of the amine and carbonyl functional groups	Other Functional Groups
amide bond	The bond between the N atom and the C atom in an amide.	Other Functional Groups
amine	An organic derivative of ammonia	Other Functional Groups

amorphous solid	A solid with no long-term structure or repetition	Solids
amphiprotic	A substance that can act as a proton donor or a proton acceptor	Brønsted-Lowry Acids and Bases
analyte	The reagent of unknown concentration	Acid-Base Titrations
angular momentum quantum number (ℓ)	An index that affects the energy and the spatial distribution of an electron in an atom.	Quantum Numbers for Electrons
anion	A species with an overall negative charge	Ions and Ionic Compounds
anode	The half cell that contains the oxidation reaction	Applications of Redox Reactions: Voltaic Cells
antibonding molecular orbital	A higher energy molecular orbital generated by destructive combination of atomic orbitals	Molecular Orbitals
aromatic hydrocarbons	Flat ring systems, which contain continuously overlapping p orbitals, such as benzene	Hydrocarbons
Arrhenius acid	A compound that increases the hydrogen ion concentration in aqueous solution	Arrhenius Acids and Bases
Arrhenius base	A compound that increases the hydroxide ion concentration in aqueous solution	Arrhenius Acids and Bases
atmosphere (atm)	A unit of pressure equal to the average atmospheric pressure at sea level; defined as exactly 760 mmHg	Pressure
atom	The smallest piece of an element that maintains the identity of that element	Atomic Theory
atomic mass	The sum of the number of protons and neutrons in a nucleus	Atomic Theory

atomic mass	The weighted average of the masses of the isotopes that compose an element	Masses of Atoms and Molecules
atomic mass unit	One-twelfth of the mass of a carbon-12 atom	Masses of Atoms and Molecules
atomic number	The number of protons in an atom	Atomic Theory
atomic radius	An indication of the size of an atom.	Periodic Trends
atomic symbol	A one- or two-letter representation of the name of an element	Atomic Theory
atomic theory	The concept that atoms play a fundamental role in chemistry	Atomic Theory
aufbau principle	The way that electrons fill the lowest energy orbitals first. From the German for “building up.”	Organization of Electrons in Atoms
autoionization constant of water (K_w)	The product of the hydrogen ion and hydroxide ion concentrations	Autoionization of Water
autoionization of water	Water molecules act as acids (proton donors) and bases (proton acceptors) with each other to a tiny extent in all aqueous solutions	Autoionization of Water
Avogadro's law	A gas law that relates number of particles to volume	Other Gas Laws
balanced chemical equation	A condition when the reactants and products of a chemical equation have the same number of atoms of all elements present	The Chemical Equation
base	A compound that increases the amount of OH^- ions in an aqueous solution	Neutralization Reactions
basic salt	An ionic compound whose aqueous solution is slightly basic	Strong and Weak Acids and Bases and Their Salts

becquerel (Bq)	A unit of radioactivity equal to 1 decay per second	Units of Radioactivity
beta particle	A type of radioactive emission equivalent to an electron	Radioactivity
boiling (or vaporization)	The process of a liquid becoming a gas	Phase Transitions: Melting, Boiling and Subliming
boiling point	The characteristic temperature at which a liquid becomes a gas	Phase Transitions: Melting, Boiling and Subliming
boiling point elevation	The increase of a solution's boiling point because of the presence of solute	Colligative Properties of Solutions
boiling point elevation constant (K_b)	The constant that relates the molality concentration of a solution and its boiling point change	Colligative Properties of Solutions
bond energy	The approximate amount of energy needed to break a covalent bond	Other Aspects of Covalent Bonding
bond order	A method of evaluating bond strength	Molecular Orbitals
bonding electron pair	A pair of electrons that makes a covalent bond	Covalent Bonds
bonding molecular orbital	The lower energy molecular orbital generated by constructive combination of atomic orbitals	Molecular Orbitals
Boyle's law	A gas law that relates pressure and volume at constant temperature and amount	Gas Laws
branched hydrocarbons	A carbon compound that is not a straight chain, having substituents appended to the longest chain	Branched Hydrocarbons

Brønsted-Lowry acid	Any species that can donate a proton to another molecule	Brønsted-Lowry Acids and Bases
Brønsted-Lowry base	Any species that can accept a proton from another molecule	Brønsted-Lowry Acids and Bases
buffer	A solution that resists dramatic changes in pH	Buffers
buffer capacity	The amount of strong acid or base a buffer can counteract	Buffers
burette or buret	A precisely calibrated volumetric delivery tube	Acid-Base Titrations
calorie	A unit of energy measurement originally defined in terms of warming up a given quantity of water. 1 cal = 4.184 J	Energy
calorimeter	A container used to measure the heat of a chemical reaction.	Enthalpy and Chemical Reactions
calorimetry	The process of measuring enthalpy changes in chemical reactions.	Enthalpy and Chemical Reactions
capillary action	The behavior of a liquid in narrow surfaces due to differences in adhesion and cohesion	Properties of Liquids
carbonyl group	A functional group where an O atom and a C atom are joined with a double bond	Other Oxygen-Containing Functional Groups
carboxyl group	A functional group composed of a carbonyl group and an OH group	Other Oxygen-Containing Functional Groups
carboxylate ion	A negatively charged ion derived from a carboxylic acid	Other Oxygen-Containing Functional Groups
carboxylic acid	A molecule with a carboxyl group	Other Oxygen-Containing Functional Groups
catalyst	A substance that increases the speed of a reaction	Shifting Equilibria: Le Chatelier's Principle

catalyst	A substance that accelerates a reaction by participating in it without being consumed	Factors that Affect the Rate of Reactions
catalyst	A substance that lowers the activation energy of a specific reaction by providing an alternate reaction pathway	Catalysis
cathode	The half cell that contains the reduction reaction	Applications of Redox Reactions: Voltaic Cells
cation	A species with an overall positive charge	Ions and Ionic Compounds
central atom	The atom in the center of a molecule	Covalent Bonds
Charles's law	A gas law that relates volume and temperature at constant pressure and amount	Gas Laws
chemical bond	The connection between two atoms in a molecule	Molecules and Chemical Nomenclature
chemical change	The process of demonstrating a chemical property	Some Basic Definitions
chemical equation	A concise way of representing a chemical reaction	The Chemical Equation
chemical equilibrium	The point at which forward and reverse chemical reactions balance each other's progress	Chemical Equilibrium
chemical nomenclature	A very specific system for naming compounds, in which unique substances get unique names	Molecules and Chemical Nomenclature
chemical property	A characteristic that describes how matter changes form in the presence of other matter	Some Basic Definitions

chemistry	The study of the interactions of matter with other matter and with energy	Introduction
coefficient	The part of a number in scientific notation that is multiplied by a power of 10	Expressing Numbers
coefficient	A number in a chemical equation indicating more than one molecule of the substance	The Chemical Equation
cohesion	The tendency of a substance to interact with itself	Properties of Liquids
colligative property	A property of solutions related to the fraction that the solute particles occupy in the solution, not their identity	Colligative Properties of Solutions
collision theory	The theory that reactions occur when reactant molecules “effectively collide”	Factors that Affect the Rate of Reactions
combined gas law	A gas law that combines pressure, volume, and temperature	Other Gas Laws
combustion reaction	A chemical reaction in which a reactant combines with oxygen to produce oxides of all other elements as products	Composition, Decomposition, and Combustion Reactions
complete ionic equation	A chemical equation in which the dissolved ionic compounds are written as separated ions	Ionic Equations: A Closer Look
composition reaction	A chemical reaction in which a single substance is produced from multiple reactants	Composition, Decomposition, and Combustion Reactions
compound	A combination of more than one element	Some Basic Definitions
compressibility factor	A measure of the extent of deviation from ideal gas behaviour.	Real Gases

concentrated solution	A solution with a lot of solute	Some Definitions
concentration	How much solute is dissolved in a given amount of solvent	Some Definitions
concentration (verb)	The removal of solvent, which increases the concentration of the solute in the solution	Dilutions and Concentrations
condensation	The process of a gas becoming a liquid	Phase Transitions: Melting, Boiling and Subliming
condensed structure	A listing of the atoms bonded to each C atom in a chain	Hydrocarbons
conjugate acid-base pair	Two species whose formulas differ by only a hydrogen ion	Brønsted-Lowry Acids and Bases
continuous spectrum	An image that contains all colours of light.	Quantum Numbers for Electrons
conversion factor	A fraction that can be used to convert a quantity from one unit to another	Converting Units
covalent bond	A chemical bond formed by two atoms sharing electrons.	Covalent Bonds
covalent network solids	A crystalline solid composed of atoms of one or more elements that are covalently bonded together in a seemingly never-ending fashion	Solids
critical point	The point at the highest temperature and pressure at which liquids and gases remain distinguishable	Properties of Liquids
crystalline solid	A solid with a regular, repeating three-dimensional structure	Solids
curie	A unit of radioactivity equal to 3.7×10^{10} decays/s	Units of Radioactivity

<i>d</i> block	The columns of the periodic table in which <i>d</i> subshells are being occupied.	Electronic Structure and the Periodic Table
Dalton's law of partial pressures	The total pressure of a gas mixture, P_{tot} , is equal to the sum of the partial pressures of the components, P_i	Gas Mixtures
daughter isotope	The product left over from the parent isotope in a nuclear equation	Radioactivity
decomposition reaction	A chemical reaction in which a single substance becomes more than one substance	Composition, Decomposition, and Combustion Reactions
degrees	The unit of temperature scales	Other Units: Temperature and Density
density	A physical property defined as a substance's mass divided by its volume	Other Units: Temperature and Density
deposition	The process of a gas becoming a solid	Phase Transitions: Melting, Boiling and Subliming
derived unit	A unit that is a product or a quotient of a fundamental unit	Expressing Units
diatomic molecule	A molecule with only two atoms	Molecules and Chemical Nomenclature
diffusion	The movement of gas molecules through one or more additional types of gas via random molecular motion	Molecular Effusion and Diffusion
dilute	A solution with very little solute	Some Definitions
dilution	The addition of solvent, which decreases the concentration of the solute in the solution	Dilutions and Concentrations

dilution equation	The mathematical formula for calculating new concentrations or volumes when a solution is diluted or concentrated	Dilutions and Concentrations
dipole-dipole interactions	An intermolecular force caused by molecules with a permanent dipole	Intermolecular Forces
dispersion force (or London dispersion force)	An intermolecular force caused by the instantaneous position of an electron in a molecule	Intermolecular Forces
dissociation	The process of an ionic compound separating into ions when it dissolves	Ionic Equations: A Closer Look
double bond	A covalent bond composed of two pairs of bonding electrons	Covalent Bonds
double-replacement reaction	A chemical reaction in which parts of two ionic compounds are exchanged	Types of Chemical Reactions: Single- and Double-Displacement Reactions
dry cell	A modern battery that does not contain large amounts of aqueous solution	Applications of Redox Reactions: Voltaic Cells
dynamic equilibrium	When a process still occurs but the opposite process also occurs at the same rate so that there is no net change in the system.	Properties of Liquids
effective nuclear charge (Z_{eff})	The net nuclear charge felt by valence electrons.	Periodic Trends
effusion	The movement of gas molecules from one container to another via a tiny hole	Molecular Effusion and Diffusion
electrodes	The cathode or anode of a voltaic cell	Applications of Redox Reactions: Voltaic Cells

electrolysis	The process of making a nonspontaneous redox reaction occur by forcing electricity into a cell	Electrolysis
electrolytic cell	A cell into which electricity is forced to make a nonspontaneous reaction occur	Electrolysis
electromagnetic spectrum	The full span of the possible wavelengths, frequencies, and energies of light.	Light
electron	A tiny subatomic particle with a negative charge	Atomic Theory
electron affinity (EA)	The energy change when a gas-phase atom accepts an electron.	Periodic Trends
electron configuration	A listing of the shell and subshell labels.	Organization of Electrons in Atoms
electron deficient molecules	A molecule with less than eight electrons in the valence shell of an atom	Violations of the Octet Rule
electron group geometry	how electron groups (bonds and nonbonding electron pairs) are arranged	Molecular Shapes and Polarity
electron groups	A covalent bond of any type or a lone electron pair	Molecular Shapes and Polarity
electron shell	A term used to describe electrons with the same principal quantum number.	Quantum Numbers for Electrons
electronegativity	A scale for judging how much atoms of any element attract electrons	Other Aspects of Covalent Bonding
electroplating	The deposition of a thin layer of metal on an object for protective or decorative purposes	Electrolysis
element	A substance that cannot be broken down into simpler chemical substances by ordinary chemical means	Some Basic Definitions

elementary step	Each event that occurs in a chemical reaction as a result of an effective collision	Reaction Mechanisms
elimination reaction	The removal of a functional group (either X or OH) and a H atom from an adjacent carbon	Alkyl halides and alcohols
endothermic	A chemical reaction that has a positive change in enthalpy.	Enthalpy and Chemical Reactions
energy	The ability to do work.	Energy
enthalpy change	The heat of a process at constant pressure. Denoted as ΔH .	Enthalpy and Chemical Reactions
enthalpy of formation	The enthalpy change for a formation reaction; denoted as ΔH_f .	Formation Reactions
enthalpy of fusion	The amount of energy needed to change from a solid to a liquid or from a liquid to a solid	Phase Transitions: Melting, Boiling and Subliming
enthalpy of sublimation	The amount of energy needed to change from a solid to a gas or from a gas to a solid	Phase Transitions: Melting, Boiling and Subliming
enthalpy of vaporization	The amount of energy needed to change from a liquid to a gas or from a gas to a liquid	Phase Transitions: Melting, Boiling and Subliming
entropy	The level of randomness (or disorder) of a system, or a measure of the energy dispersal of the molecules in the system	Entropy and the Second Law of Thermodynamics
enzyme	Protein molecules which serve to catalyze biochemical reactions	Catalysis
enzyme-substrate complex	The binding of substrate to the enzymatic active site	Catalysis
equilibrium constant (K_{eq})	A numerical value that relates to the ratio of products and reactants at equilibrium	The Equilibrium Constant

equivalence point	The point of the reaction when all the analyte has been reacted with the titrant	Acid-Base Titrations
ester group	A functional group made by combining a carboxylic acid with an alcohol	Other Oxygen-Containing Functional Groups
ether group	A functional group that has an O atom attached to two organic groups	Other Oxygen-Containing Functional Groups
evaporation	The formation of a gas phase from a liquid at temperatures below the boiling point	Properties of Liquids
exact number	A number from a defined relationship that technically has an infinite number of significant figures	Converting Units
exothermic	A chemical reaction that has a negative change in enthalpy.	Enthalpy and Chemical Reactions
expanded valence shell molecules	A molecule with more than eight electrons in the valence shell of an atom	Violations of the Octet Rule
experiment	A test of the natural universe to see if a guess (hypothesis) is correct	Chemistry as a Science
exponent	The raised number to the right of a 10 indicating the number of factors of 10 in the original number	Expressing Numbers
<i>f</i> block	The columns of the periodic table in which <i>f</i> subshells are being occupied.	Electronic Structure and the Periodic Table
fission	The breaking apart of an atomic nucleus into smaller nuclei	Radioactivity
formation reaction	A chemical reaction that forms one mole of a substance from its constituent elements in their standard states.	Formation Reactions

freezing point depression	The decrease of a solution's freezing point because of the presence of solute	Colligative Properties of Solutions
freezing point depression constant (K _f)	The constant that relates the molality concentration of a solution and its freezing point change	Colligative Properties of Solutions
frequency	The number of cycles of light that pass a given point in one second.	Light
frequency factor (A)	A factor that takes into account the frequency of reactions and the likelihood of correct molecular orientation	Activation Energy and the Arrhenius Equation
frontier molecular orbitals	A term which refers to the HOMO and LUMO, the most likely orbitals to be involved in chemical reactions or processes	Molecular Orbitals
functional group	A collection of atoms or bonds with certain characteristic reactions	Alkyl halides and alcohols
fundamental units	One of the seven basic units of SI used in science	Expressing Units
gamma ray	A type of radioactive emission that is a very energetic form of electromagnetic radiation	Radioactivity
gas law	A simple mathematical formula that allows one to model, or predict, the behaviour of a gas	Gas Laws
Gay-Lussac's law	A gas law that relates pressure with absolute temperature	Other Gas Laws
Geiger counter	An electrical device that detects radioactivity	Units of Radioactivity
Gibbs free energy (G)	A measure of spontaneity which incorporates both enthalpy and entropy	Gibbs Free Energy

Graham's law of effusion	A law that relates the rate of effusion of a gas to the inverse of the square root of its molar mass.	Molecular Effusion and Diffusion
gray (Gy)	A unit of radioactive exposure equal to 100 rad	Units of Radioactivity
half cell	A part of a voltaic cell that contains one half reaction	Applications of Redox Reactions: Voltaic Cells
half reaction	The individual oxidation or reduction reaction of a redox reaction	Balancing Redox Reactions
half reaction method	The method of balancing redox reactions by writing and balancing the individual half reactions	Balancing Redox Reactions
half-life	The amount of time it takes for one-half of a radioactive isotope to decay	Half-Life
half-life	The amount of time required for the concentration of a reactant to drop to one half of its initial concentration	Concentration-Time Relationships: Integrated Rate Laws
heat	The transfer of energy from one body to another due to a difference in temperature.	Work and Heat
heating curve	A plot of the temperature versus the amount of heat added	Phase Transitions: Melting, Boiling and Subliming
Hess's law	When chemical equations are combined algebraically, their enthalpies can be combined in exactly the same way.	Hess's Law
heterogeneous catalyst	A catalyst that is in a different phase from one or more of the reactants	Catalysis

heterogeneous equilibrium	An equilibrium in which more than one phase of reactants or products is present	The Equilibrium Constant
heterogeneous mixture	A non-uniform combination of more than one substance	Some Basic Definitions
HOMO	The highest occupied molecular orbital	Molecular Orbitals
homogeneous catalyst	A catalyst that is present in the same phase as the reactant molecules	Catalysis
homogeneous mixture	A uniform mixture of more than one substance that behaves as a single substance	Some Basic Definitions
Hund's rule	One electron is placed in each degenerate orbital before pairing electrons in the same orbital.	Organization of Electrons in Atoms
hybridization	A mathematical mixing of atomic orbitals	Valence Bond Theory and Hybrid Orbitals
hydrocarbons	An organic compound composed of carbon and hydrogen	Hydrocarbons
hydrogen bonding	The very strong interaction between molecules due to H atoms being bonded to N, O, or F atoms	Intermolecular Forces
hydrogenation reaction	The reaction of hydrogen across a C–C double or triple bond, usually in the presence of a catalyst	Hydrocarbons
hydronium ion	The actual chemical species that represents a hydrogen ion in aqueous solution	Arrhenius Acids and Bases
hypothesis	An educated guess about how the natural universe works	Chemistry as a Science
hydrolysis	A reaction with water	Brønsted-Lowry Acids and Bases

ICE chart	A table used to calculate equilibria values featuring rows of initial, change and equilibria concentration	Calculating Equilibrium Constant Values
ideal gas	A gas that conforms exactly to the tenets of the kinetic molecular theory.	Real Gases
ideal gas law	A gas law that relates all four independent physical properties of a gas under any conditions	The Ideal Gas Law and Some Applications
indicator	A substance whose color change indicates the equivalence point of a titration	Acid-Base Titrations
initial rate	The instantaneous rate at the start of a reaction	Reaction Rates
initial rates method	A method to determine the rate law from the instantaneous reaction rate upon mixing the reactants	Rate Laws
instantaneous reaction rate	The rate of reaction at one instant in time	Reaction Rates
intermediate	A chemical species does not appear in the overall balanced equation and is generated in one elementary step but used up in a subsequent step	Reaction Mechanisms
ion	A species with an overall electric charge	Ions and Ionic Compounds
ionic compound	A compound formed from positive and negative ions	Ions and Ionic Compounds
ionic formula	The chemical formula for an ionic compound	Ions and Ionic Compounds
ionic solid	A crystalline solid composed of ions	Solids
ionization energy (IE)	The amount of energy required to remove an electron from an atom in the gas phase.	Periodic Trends

isolated system	A system that does not allow a transfer of energy or matter into or out of itself.	Energy
isomer	A molecule with the same molecular formula as another molecule but a different structure	Hydrocarbons
isothermal	A process that does not change the temperature	Phase Transitions: Melting, Boiling and Subliming
isotopes	Atoms of the same element that have different numbers of neutrons	Atomic Theory
joule	The SI unit of energy.	Energy
kelvin	The fundamental unit of temperature in SI	Other Units: Temperature and Density
ketone	A compound where the carbonyl carbon is attached to two carbon chains	Other Oxygen-Containing Functional Groups
kinetic energy	Energy due to motion	Kinetic-Molecular Theory of Gases
kinetic molecular theory of gases	A model that helps us understand gases at the molecular level and their physical properties	Kinetic-Molecular Theory of Gases
kinetics	The study of reaction rate and the factors that can influence reaction rate	Introduction to Kinetics
law of conservation of energy	Law of physics that states that the total energy of an isolated system does not increase or decrease.	Energy
law of mass action	The relationship of the amounts of reactants and products at equilibrium	The Equilibrium Constant
Le Chatelier's principle	If an equilibrium is stressed, then the reaction shifts to reduce the stress	Shifting Equilibria: Le Chatelier's Principle

Lewis diagram	A representation of the valence electrons of an atom that uses dots around the symbol of the element.	Lewis Electron Dot Diagrams
limiting reagent	The reactant that runs out first for a given chemical reaction	Limiting Reagents
line spectrum	An image that contains only certain colors of light	Quantum Numbers for Electrons
locant	The numerical position of a substituent	Branched Hydrocarbons
lock and key model	A simple model used to describe enzyme activity, where substrates must fit into appropriately shaped active sites	Catalysis
lone electron pairs	A pair of electrons that does not make a covalent bond	Covalent Bonds
LUMO	The lowest unoccupied molecular orbital	Molecular Orbitals
magnetic quantum number (m_l)	The index that determines the orientation of the electron's spatial distribution.	Quantum Numbers for Electrons
mass-mass calculation	A calculation in which you start with a given mass of a substance and calculate the mass of another substance involved in the chemical equation	Mole-Mass and Mass-Mass Calculations
matter	Anything that has mass and takes up space.	Some Basic Definitions
mean free path	The average distance travelled by a molecule between collisions.	Molecular Effusion and Diffusion
melting	The process of a solid becoming a liquid	Phase Transitions: Melting, Boiling and Subliming
melting point	The characteristic temperature at which a solid becomes a liquid	Phase Transitions: Melting, Boiling and Subliming

meniscus	The curved surface a liquid makes as it approaches a solid barrier	Properties of Liquids
metal	An element that conducts electricity and heat well and is shiny, silvery, solid, ductile, and malleable	Some Basic Definitions
metallic solid	A solid with the characteristic properties of a metal	Solids
microstate (W)	A term used to describe different possible arrangements of molecular position and kinetic energy, at a particular thermodynamic state	Entropy and the Second Law of Thermodynamics
millimeters of mercury (mmHg)	The amount of pressure exerted by a column of mercury exactly 1 mm high	Pressure
mixture	A physical combination of more than one substance	Some Basic Definitions
molality (m)	The number of moles of solute per kilogram of solvent	Quantitative Units of Concentration
molar mass	The mass of 1 mol of a substance in grams	The Mole
molar volume	The volume of exactly 1 mol of a gas; equal to 22.4 L at STP.	The Ideal Gas Law and Some Applications
molarity (M)	The number of moles of solute divided by the number of liters of solution	Quantitative Units of Concentration
mole	The number of things equal to the number of atoms in exactly 12 g of carbon-12; equals 6.022×10^{23} things	The Mole
mole fraction	The ratio of the number of moles of a component in a mixture divided by the total number of moles in the sample	Gas Mixtures

mole fraction	The ratio of the number of moles of a component to the total number of moles in a system	Colligative Properties of Solutions
molecular formula	A formal listing of what and how many atoms are in a molecule	Molecules and Chemical Nomenclature
molecular geometry	how the atoms in a molecule are arranged	Molecular Shapes and Polarity
molecular mass	The sum of the masses of the atoms in a molecule	Masses of Atoms and Molecules
molecular orbital theory (MO theory)	A more sophisticated model of chemical bonding where new molecular orbitals are generated using a mathematical process called Linear Combination of Atomic Orbitals (LCAO)	Molecular Orbitals
molecular polarity	The vector sum of the individual bond dipoles	Molecular Shapes and Polarity
molecular solid	A crystalline solid whose components are covalently bonded molecules	Solids
molecularity	The total number of molecules that participate in the effective collision of the elementary step	Reaction Mechanisms
molecule	The smallest part of a substance that has the physical and chemical properties of that substance	Molecules and Chemical Nomenclature
mole-mass calculation	A calculation in which you start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa	Mole-Mass and Mass-Mass Calculations

mole-mole calculation	A stoichiometry calculation when one starts with moles of one substance and convert to moles of another substance using the balanced chemical equation	The Mole in Chemical Reactions
monomer	The repeated unit of a polymer	Polymers
net ionic equation	A chemical equation with the spectator ions removed	Ionic Equations: A Closer Look
neutral salt	An ionic compound that does not affect the acidity of its aqueous solution	Strong and Weak Acids and Bases and Their Salts
neutralization reaction	The reaction of an acid with a base to produce water and a salt	Neutralization Reactions
neutron	A subatomic particle with no charge	Atomic Theory
node (nodal plane)	An area of zero electron density	Molecular Orbitals
nomenclature	The rules of naming in organic chemistry	Branched Hydrocarbons
nonmetal	An element that exists in various colors and phases, is brittle, and does not conduct electricity or heat well	Some Basic Definitions
nonpolar covalent bond	The equal sharing of electrons in a covalent bond	Other Aspects of Covalent Bonding
normal boiling point	The characteristic temperature at which a liquid becomes a gas when the surrounding pressure is exactly 1 atm	Phase Transitions: Melting, Boiling and Subliming
nuclear energy	The controlled harvesting of energy from fission reactions	Nuclear Energy
nuclear equation	A chemical equation that emphasizes changes in atomic nuclei	Radioactivity

nuclear model	The model of an atom that has the protons and neutrons in a central nucleus with the electrons in orbit about the nucleus	Atomic Theory
nucleus	The centre of an atom that contains protons and neutrons	Atomic Theory
odd-electron molecules	A molecule with an odd number of electrons in the valence shell of an atom	Violations of the Octet Rule
orbital	The specific set of principal, angular momentum, and magnetic quantum numbers for an electron.	Quantum Numbers for Electrons
osmosis	The tendency of solvent molecules to pass through a semipermeable membrane due to concentration differences	Colligative Properties of Solutions
osmotic pressure	The tendency of a solution to pass solvent through a semipermeable membrane due to concentration differences	Colligative Properties of Solutions
oxidation	The loss of one or more electrons by an atom; an increase in oxidation number	Oxidation-Reduction Reactions
oxidation number	A number assigned to an atom that helps keep track of the number of electrons on the atom	Oxidation-Reduction Reactions
oxidation-reduction (redox) reactions	A chemical reaction that involves the transfer of electrons	Oxidation-Reduction Reactions
p block	The columns of the periodic table in which p subshells are being occupied.	Electronic Structure and the Periodic Table
parent isotope	The reactant in a nuclear equation	Radioactivity
parts per billion (ppb)	Ratio of mass of solute to total mass of sample times 1,000,000,000	Quantitative Units of Concentration

parts per million (ppm)	Ratio of mass of solute to total mass of sample times 1,000,000	Quantitative Units of Concentration
parts per thousand (ppth)	Ratio of mass of solute to total mass of sample times 1,000	Quantitative Units of Concentration
Pauli exclusion principle	No two electrons in an atom can have the same set of four quantum numbers.	Organization of Electrons in Atoms
percent yield	Actual yield divided by theoretical yield times 100% to give a percentage between 0% and 100%	Yields
percentage composition by mass (or mass percentage, % m/m)	Ratio of mass of solute to the total mass of a sample times 100	Quantitative Units of Concentration
periodic table	A chart of all the elements	Atomic Theory
periodic trends	The variation of properties versus position on the periodic table.	Periodic Trends
pH	The negative logarithm of the hydrogen ion concentration	The pH Scale
pH scale	The range of values from 0 to 14 that describes the acidity or basicity of a solution	The pH Scale
phase	An important physical property that defines whether matter is a solid, liquid, gas or supercritical fluid	Some Basic Definitions
phase diagram	A graphical representation of the equilibrium relationships that exist between the phases of a substance under specified pressures and temperatures	Properties of Liquids
photon	The name of a discrete unit of light acting as a particle.	Light

physical change	A change that occurs when a sample of matter changes one or more of its physical properties	Some Basic Definitions
physical property	A characteristic that describes matter as it exists	Some Basic Definitions
pi bond (π bond)	The sideways overlap of p orbitals, placing electron density on opposite sides of the inter-nuclear axis – a double or triple bond	Valence Bond Theory and Hybrid Orbitals
Planck's constant	The proportionality constant between the frequency and the energy of light: 6.626×10^{-34} J·s.	Light
pOH	The negative logarithm of the hydroxide ion concentration	The pH Scale
polar covalent bond	A covalent bond between different atoms that attract the shared electrons by different amounts and cause an imbalance of electron distribution	Other Aspects of Covalent Bonding
polarity	A measure of the unequal sharing of electrons which has resulted in a dipole moment	Other Aspects of Covalent Bonding
polyatomic ions	An ion that contains more than one atom	Ions and Ionic Compounds
polymer	A long molecule made of many repeating units	Polymers
polymerization	The process of making a polymer	Polymers
polyprotic acid	An acid capable of donating more than one H^+ ion	Some Special Types of Equilibria
precipitate	A solid that falls out of solution in a precipitation reaction	Types of Chemical Reactions: Single- and Double-Displacement Reactions

precipitation reaction	A chemical reaction in which two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve	Types of Chemical Reactions: Single- and Double-Displacement Reactions
prefix	A prefix used with a unit that refers to a multiple or fraction of a fundamental unit to make a more conveniently sized unit for a specific quantity	Expressing Units
pressure	Force per unit area	Pressure
primary battery	A battery that cannot be recharged	Applications of Redox Reactions: Voltaic Cells
principal quantum number (n)	The index that largely determines the energy of an electron in an atom.	Quantum Numbers for Electrons
product	A final substance in a chemical equation	The Chemical Equation
proton	A subatomic particle with a positive charge	Atomic Theory
qualitative	A description of the quality of an object	Chemistry as a Science
quantitative	A description of a specific amount of something	Chemistry as a Science
quantization	When a quantity is restricted to having only certain values.	Quantum Numbers for Electrons
quantum mechanics	The theory of electrons that treats them as a wave.	Quantum Numbers for Electrons
quantum number	An index that corresponds to a property of an electron, like its energy.	Quantum Numbers for Electrons
rad	A unit of radioactive exposure equal to 0.01 J/g of tissue	Units of Radioactivity
radioactive decay	The spontaneous change of a nucleus from one element to another	Radioactivity

radioactivity	Emanations of particles and radiation from atomic nuclei	Radioactivity
Raoult's law	The mathematical formula for calculating the vapor pressure of a solution	Colligative Properties of Solutions
rate constant (k)	A proportionality constant specific to each reaction at a particular temperature	Rate Laws
rate-determining step	The slowest step in a multistep mechanism	Reaction Mechanisms
rate law	A mathematical relationship between the reaction rate and the reactant concentrations	Rate Laws
reactant	An initial substance in a chemical equation	The Chemical Equation
reaction mechanism	The bond making and bond breaking steps which occur at the molecular level during a chemical reaction	Reaction Mechanisms
reaction order	The sum of the concentration term exponents in a rate law equation	Rate Laws
reaction rate	The speed of a chemical reaction	Introduction to Kinetics
real gas	A gas that deviates from ideal behaviour.	Real Gases
redox reaction	A chemical reaction that involves the transfer of electrons	Oxidation-Reduction Reactions
reduction	The gain of one or more electrons by an atom; a decrease in oxidation number	Oxidation-Reduction Reactions
rem	A unit of radioactive exposure that includes a factor to account for the type of radioactivity	Units of Radioactivity

ribozyme	Ribonucleic acid (RNA) molecules capable of catalyzing certain chemical reactions	Catalysis
root-mean-square (rms) speed (urms)	The speed of molecules having exactly the same kinetic energy as the average kinetic energy of the sample	Kinetic-Molecular Theory of Gases
s block	The columns of the periodic table in which s subshells are being occupied	Electronic Structure and the Periodic Table
salt	Any ionic compound that is formed from a reaction between an acid and a base	Neutralization Reactions
salt	Any ionic compound that is formed from a reaction between an acid and a base	Arrhenius Acids and Bases
salt bridge	A part of a voltaic cell that contains a solution of some ionic compound whose ions migrate to either side of the voltaic cell to maintain the charge balance	Applications of Redox Reactions: Voltaic Cells
saturated hydrocarbons	A carbon compound with the maximum possible number of H atoms in its formula	Hydrocarbons
saturated solution	A solution with the maximum amount of solute dissolved in it	Some Definitions
science	The process of knowing about the natural universe through observation and experiment	Chemistry as a Science
scientific law	A specific statement that is thought to be never violated by the entire natural universe	Chemistry as a Science
scientific notation	An expression of a number using powers of 10	Expressing Numbers

screening	The repelling valence electrons by core electrons	Periodic Trends
second law of thermodynamics	A spontaneous process will increase the entropy of the universe	Entropy and the Second Law of Thermodynamics
secondary battery	A battery that can be recharged	Applications of Redox Reactions: Voltaic Cells
semimetal	An element that has properties of both metals and nonmetals	Some Basic Definitions
semipermeable membrane	A thin membrane that will pass certain small molecules but not others	Colligative Properties of Solutions
SI unit	International System of Units used by all scientists, literally translated from “le Système International d’unités.”	Expressing Units
Sievert (Sv)	Sievert (Sv) is a related unit and is defined as 100 rem	Units of Radioactivity
sigma bond (σ bond)	Orbital overlap to form a bond which has cylindrical symmetry – a single bond	Valence Bond Theory and Hybrid Orbitals
significant figures	The limit of the number of places a measurement can be properly expressed with	Significant Figures
silicones	A polymer based on a silicon and oxygen backbone	Polymers
single bond	A covalent bond composed of one pair of electrons	Covalent Bonds
single-replacement reaction	A chemical reaction in which one element is substituted for another element in a compound	Types of Chemical Reactions: Single- and Double-Displacement Reactions
solidification	The process of a liquid becoming a solid	Phase Transitions: Melting, Boiling and Subliming

solubility	The maximum amount of a solute that can be dissolved in a given amount of a solvent	Some Definitions
solubility rules	General statements that predict which ionic compounds dissolve and which do not	Types of Chemical Reactions: Single- and Double-Displacement Reactions
solute	The minor component of a solution	Some Definitions
solution	See homogeneous mixture	Some Basic Definitions
solvent	The major component of a solution.	Some Definitions
specific heat capacity	The proportionality constant between heat, mass, and temperature change; also called specific heat.	Work and Heat
spectator ion	An ion that does nothing in the overall course of a chemical reaction	Ionic Equations: A Closer Look
spin quantum number (m_s)	The index that indicates one of two spin states for an electron.	Quantum Numbers for Electrons
spontaneous process	A process that occurs without the influence of external forces or a change that moves a system towards equilibrium	Spontaneous Change
standard molar entropy (S°)	The entropy of 1 mole of a substance in its standard state, at 1 atm of pressure	Measuring Entropy and Entropy Changes
standard notation	A straightforward expression of a number	Expressing Numbers
standard temperature and pressure (STP)	A set of benchmark conditions used to compare other properties of gases: 100 kPa for pressure and 273 K for temperature	The Ideal Gas Law and Some Applications

stoichiometry	The relating of one chemical substance to another using a balanced chemical reaction	Stoichiometry
strong acid	Any acid that is 100% dissociated into ions in aqueous solution	Strong and Weak Acids and Bases and Their Salts
strong base	Any base that is 100% dissociated into ions in aqueous solution	Strong and Weak Acids and Bases and Their Salts
sublimation	The process of a solid becoming a gas	Phase Transitions: Melting, Boiling and Subliming
subshell	A term used to describe electrons in a shell that have the same angular momentum quantum number.	Quantum Numbers for Electrons
substance	Matter that has the same physical and chemical properties throughout.	Some Basic Definitions
substituent	A branch off a main chain in a hydrocarbon	Branched Hydrocarbons
substrate	The reactants which are specific for a biological catalyst	Catalysis
supercritical fluid	A phase beyond the critical point, where liquid and gas phases are no longer distinct	Properties of Liquids
supersaturated solution	A unstable solution with more than the normal maximum amount of solute in it	Some Definitions
surface tension	An effect caused by an imbalance of forces on the atoms at the surface of a liquid	Properties of Liquids
surrounding atoms	An atom that makes covalent bonds to the central atom(s)	Covalent Bonds
system	The part of the universe that is under study.	Energy

temperature	A measure of the average amount of kinetic energy a system contains	Other Units: Temperature and Density
theoretical yield	An amount that is theoretically produced as calculated using the balanced chemical reaction	Yields
theory	A general statement that explains a large number of observations	Chemistry as a Science
thermochemical equation	A chemical equation that includes an enthalpy change.	Enthalpy and Chemical Reactions
thiol	The sulfur analog of an alcohol	Other Functional Groups
third law of thermodynamics	At absolute zero the entropy of a pure, perfect crystal is zero	Measuring Entropy and Entropy Changes
titrant	The reagent of known concentration	Acid-Base Titrations
titration	A chemical reaction performed quantitatively to determine the exact amount of a reagent	Acid-Base Titrations
torr	Another name for a millimeter of mercury	Pressure
tracer	A substance that can be used to follow the pathway of that substance through a structure	Uses of Radioactive Isotopes
transition state	The highest energy transitional point in the elementary step	Reaction Mechanisms
triple bond	A covalent bond composed of three pairs of bonding electrons	Covalent Bonds
unsaturated hydrocarbons	A carbon compound with less than the maximum possible number of H atoms in its formula	Hydrocarbons
unsaturated solution	A solution with less than the maximum amount of solute dissolved in it	Some Definitions

valence electron	An electron in the highest-numbered shell or in the last unfilled subshell. Valence electrons are those that are most likely to be involved in chemical reactions.	Electronic Structure and the Periodic Table
valence shell	The highest-numbered shell in an atom that contains electrons.	Electronic Structure and the Periodic Table
valence shell electron pair repulsion theory (VSEPR)	The general concept that estimates the shape of a simple molecule: electron pairs repel each other to get as far away from each other as possible	Molecular Shapes and Polarity
van der Waals equation	An equation that compensates for deviations from ideal gas behaviour, correcting for intermolecular forces and the volume of gas molecules.	Real Gases
van't Hoff factor (i)	The number of particles each solute formula unit breaks apart into when it dissolves	Colligative Properties of Ionic Solutes
vapor	Material in the gas phase due to evaporation	Properties of Liquids
vapor pressure	The partial pressure exerted by evaporation of a liquid	Gas Mixtures
vapor pressure depression	The decrease of a solution's vapor pressure because of the presence of a solute	Colligative Properties of Solutions
vector quantity	A quantity which has both a magnitude and direction	Molecular Shapes and Polarity
voltaic (galvanic) cell	An apparatus that allows for useful electrical work to be extracted from a redox reaction.	Applications of Redox Reactions: Voltaic Cells
wavelength	The distance between corresponding points in two adjacent light cycles.	Light

weak acid	Any acid that is less than 100% dissociated into ions in aqueous solution	Strong and Weak Acids and Bases and Their Salts
weak base	Any base that is less than 100% dissociated into ions in aqueous solution	Strong and Weak Acids and Bases and Their Salts

This page provides a record of edits and changes made to this book since its initial publication. Whenever edits or updates are made to the text, we provide a record and description of those changes here. If you have a correction or recommendation you would like to suggest, please contact the Study Support team at **studysupport@rmit.edu.au**.

Version	Date	Change	Details
Pre-publication	15th December 2023		Internal review completed, ready for external review.