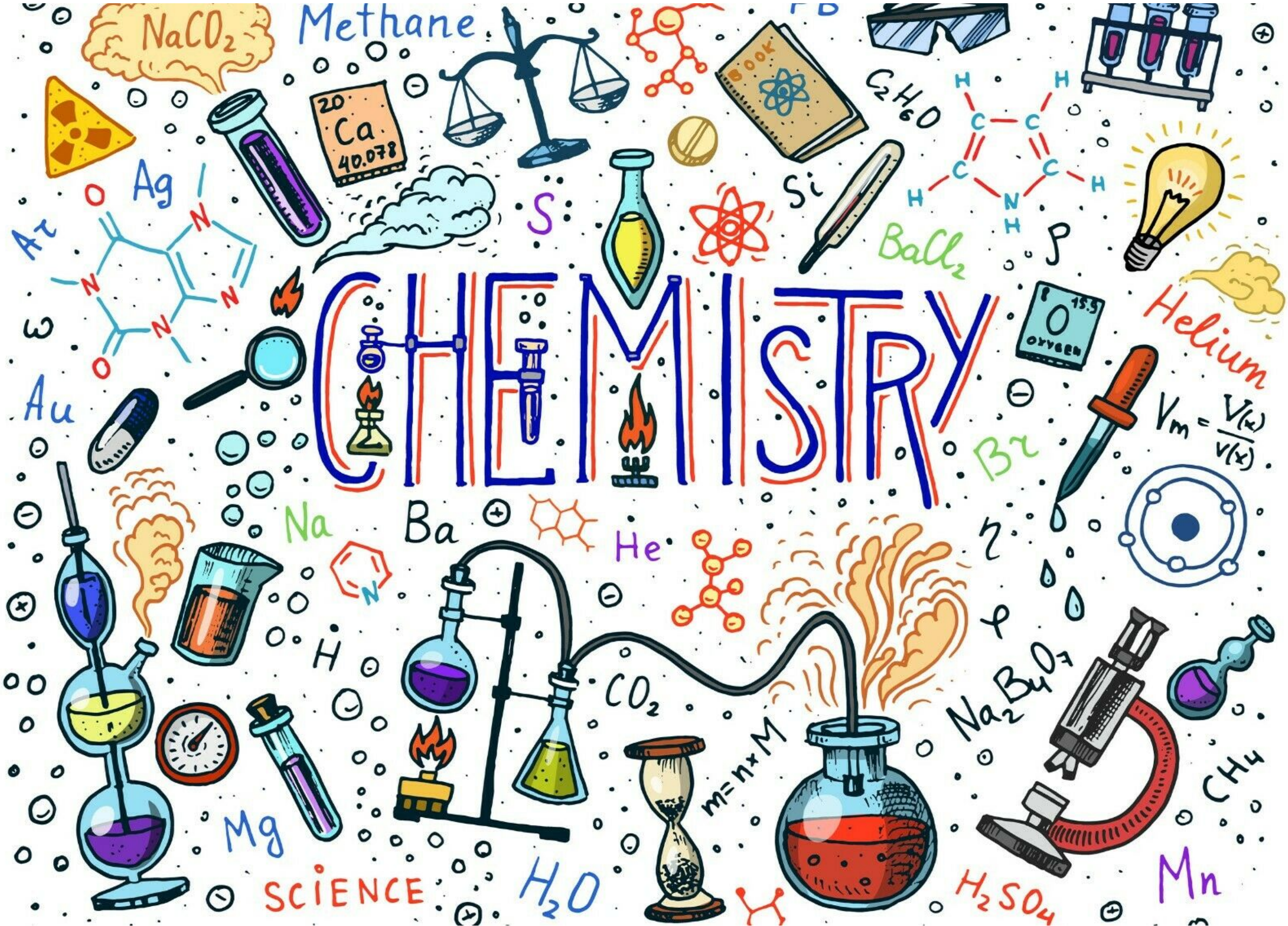
NEEHAR ARJUNWADKAR

WITH SURABHEE & MIHIR ARJUNWADKAR

CHEMISTRY IN THE TIME OF LOCKDOWN



Dedicated to all those who enjoy chemistry

— and exploring things with their own hands!

Released February 28, 2021 :: Indian National Science Day (W National_Science_Day)

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Foreword: Chemical adventures of a curious character

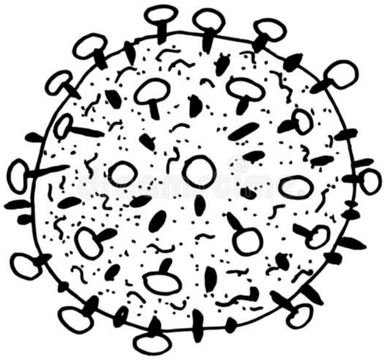

Figure 1: A novel caricature of the 2020 novel coronavirus.

This little book is about the author's explorations in chemistry during the 2020 COVID19 lockdown. My own role in his chemistry adventures has been that of a slightly bedazzled bystander cum supplier of chemicals cum helper cum videographer cum cleaner. Of late, my job description has changed into "writing assistance, full editorial support, copy editing, and dealing with the idiosyncrasies of LaTeX". Apart from many exciting discussions, my role in this book is limited to these roles — except for the last chapter which is written jointly by us, the side-notes, some compulsive edits on my part – giving a mixed flavour to the writing style. The chemistry ideas and the execution/explorations described here are uniquely his.

According to one knowledgeable biographical source[1], the author's fascination for chemistry can be traced to an early age when he watched a video[2] which showed the explosion of a fish tank full of water when a piece of rubidium was dropped in it[3]. About the same time, he also watched a video clip of a feather gently landing on freshly synthesized nitrogen tri-iodide exploding into purple-brown iodine smoke[4]. While his fascination with explosions and fire continues in his teen years, of late, he has grown into a careful chemistry explorer who makes sure that neither he nor anybody else gets hurt in his experiments.

For the explorations described here, he has mostly used chemicals which are available off-the-shelf in medical shops

[1] Yours truly.

[2] Such as ▶ OYNslaSbFdg

[3] No animal seemed to have been harmed. There were no fish nor other aquatic life in the tank. Do NOT try this at home or outdoors – even if you can get hold of any of the alkali metals in any amount.

[4] Such as ▶ DFfRqoldArM

(glycerine/glycerol, boric acid, potassium permanganate, 3% or 6% hydrogen peroxide, ethanol or isopropyl alcohol, and acetone), in the home cleaning supplies shops (30% hydrochloric acid), and in the kitchen (white vinegar/acetic acid, aluminium foil, lemon juice, baking soda, and egg shells as a source of calcium carbonate). Very occasionally, he has toyed with industry- or agriculture-grade chemicals available off-the-shelf in small quantities (sodium hydroxide, acetone, copper sulphate, potassium nitrate, and isopropyl alcohol). As lab equipment, he has used a few glass test tubes, kitchen aluminium foil, small containers improvised from soft-drink cans, well-cleaned[5] plastic or glass bottles, plastic straw, toothpicks, live or burnt matchsticks, disposable spoons and cups, and disposable syringes[6].

In this book, he makes no attempt at completeness. First and foremost, he enjoys and relishes his hands-on explorations[7]. Then, he goes on to seek explanations: Explanations may come from his own ever-evolving understanding of chemistry, speculation, discussions, more explorations, or through what can be scavenged easily from friendly internet sources. In any case, all explanations – correct or incorrect – are taken as *tentatively correct until contradicted* – by observation or better understanding.

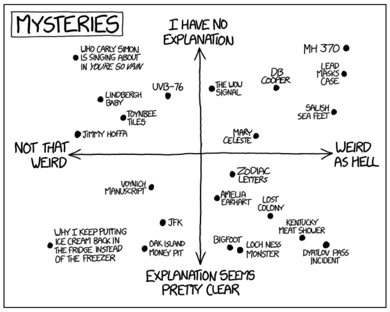

Figure 2: XKCD's view on mysteries and explanations.

References and information sources are cited in this book in the form of abbreviated links to Wikipedia, YouTube and such other sources, usually as side notes. In the PDF version of this book, these links are clickable. If you are somehow reading a print version of this book, then the prescription for constructing the full URLs is provided at the end of this foreword: You will need to carefully type the full URL into your web browser and then hit ENTER[8].

Images in this book which are taken from internet and other sources are gratefully acknowledged at the end of the book. In the PDF version, these images are also linked to their respective internet sources. Please be assured that this book is the

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[5] Different people usually have widely/wildly different standards for cleanliness.

[6] After safely discarding their needles.

[7] The scene is much like the Druids' Conference in Asterix & the Goths.

[8] QR codes would have been great, but they are a bit too bulky considering the number of links mentioned in this book.

result of a fun, educational, non-commercial project during the 2020 COVID19 lockdown, and we don't mean to infringe upon anybody's rights, intellectual or otherwise.

Special thanks are due to two friends for enthusiastically encouraging this little book project. One is Jayant Gadgil, chemist and artist. The other one is Niruj Mohan, astrophysicist, whose deep passion for science outreach has, in part, inspired this book project. Special thanks also to the creators of TeX, LaTeX, the `tuftebook` document class, and the `R` computing platform / programming language. Last, but not the least, special thanks to Wikipedia and such other information resources, various image sources, creative creators of chemistry videos on YouTube and elsewhere, and so on.

If you notice any errors, typos, etc., please let us know at the email address mentioned on page 2. We will be happy to make corrections in the next revision.

Enjoy this book and the explorations described here. Adult supervision is strongly recommended. Be safe, enjoy, have fun!

Mihir Arjunwadkar
Pune, India
28 February 2021
W National_Science_Day

How to construct full URLs from abbreviations

URL KEY	FULL URL
▶ OYNslaSbFdg	`https://youtu.be/OYNslaSbFdg`
W Chemical_reaction	`https://en.wikipedia.org/wiki/Chemical_reaction`
Chemiday 3-1-0-356	`https://chemiday.com/en/reaction/3-1-0-356`

Preface

Chemistry is a fascinating science. Chemistry allows us to understand the world around us. Everything that we touch, smell, taste and feel is because of chemical processes in the body. Chemistry lies in between physics and biology. Chemistry explores and studies a rich world of phenomena where atoms and molecules interact with one another.

> *Imagine putting a piece of metal in a liquid that looks like water ... Its starts to bubble, gets hot, a gas comes out which burns with a* POP *... the metal vanishes ... the color of the solution changes suddenly!*
>
> *Imagine putting a spoonful of a white powder in a washing machine ... and your clothes come out sparkling clean!*

This is why I find chemistry fascinating!

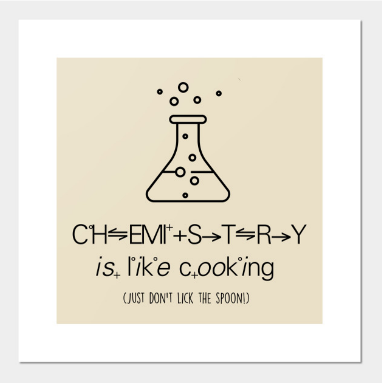

Figure 3: "Chemistry is like cooking (just dont lick the spoon)"!

Neehar Arjunwadkar
Pune, India
28 February 2021
W National_Science_Day

Elephant's toothpaste

YouTube has many videos on making "elephant's toothpaste". If you have searched for it then you know how many videos there are and what the experiment is. But if you haven't, here's a quick description of the experiment: Elephant's toothpaste is an experiment in which there is a liquid mixture in one container. Another solution is then poured in this container. As soon as the two solutions mix, they start forming some sort of foam almost instantly. The foam can even spout out of the container upwards a few meters in height before falling down. If any of the two solutions is colored, then the foam takes on that color. Here is one video of this experiment: `https://sciencebob.com/fantastic-foamy-fountain/`.

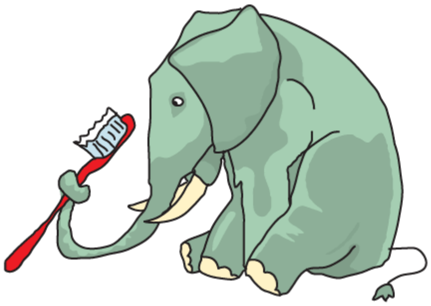

Figure 4: To brush or not to brush – an elephant's daily dilemma.

Figure 5: Elephant's toothpaste rising and spilling out of a conical flask.

The chemistry of elephant's toothpaste

What happens here is a *chemical reaction*[1]. A chemical reaction turns *reactants* into *products*, sometimes with the help of a *catalyst*[2]. A catalyst is a chemical substance which does not take part the reaction itself, but it helps the reaction happen, or happen faster. We will talk about a chemical reaction in more detail in another chapter later on.

In the elephant's toothpaste experiment, the first solution is made of hydrogen peroxide[3], soap and color. A hydrogen peroxide (H_2O_2) molecule is like a water[4] molecule (H_2O), but with an extra oxygen[5] atom in between. The first solution is

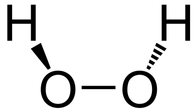

Figure 6: Hydrogen peroxide molecule.

[1] W Chemical_reaction

[2] W Catalysis

[3] W Hydrogen_peroxide

[4] W Water

[5] W Oxygen

kept in a conical flask (see Figure 5) or a plastic bottle with a narrow neck. The second solution is made by adding a catalyst like potassium iodide (KI)[6] to water, plus some color if you like. It doesn't matter much which solution has the color.

The basic chemical reaction which takes place here is[7]

$$2\,H_2O_2 \longrightarrow 2\,H_2O + O_2 + \text{Heat}. \qquad (1)$$

In this reaction, two hydrogen peroxide molecules decompose into two water molecules and one oxygen molecule. Water is a liquid, but oxygen is a gas – so it tries to escape. On the way out, it is captured by the soap mixed in the solution. This is what forms the foam. The foam is hot because a lot of heat[8] is released by the decomposition reaction. Sometimes, the rising foam catches the still-decomposing H_2O_2.This makes it swell even more – you can notice this after the foam falls on the table top.

As potassium iodide (KI) is not that easily available, you can use potassium permanganate ($KMnO_4$; see Fig. 8 and W Potassium_permanganate) instead of potassium iodide. $KMnO_4$ is not a catalyst, but in this experiment, it gives similar result as KI. We will explore the chemical reaction between H_2O_2 and $KMnO_4$ in the next chapter. This reaction (Reaction (2)) too generates oxygen and it forms foam – that is, elephant's toothpaste – because of the soap in the solution. That this gas is indeed oxygen is verified in another experiment – this is described in the next chapter.

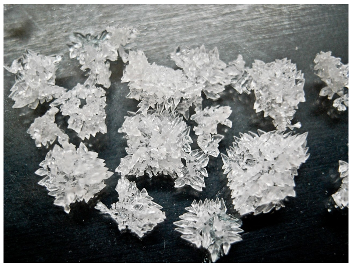

Figure 7: Potassium iodide (KI) crystals.

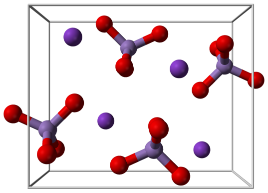

Figure 8: Ball-and-stick model of the $KMnO_4$ crystal structure; light purple: manganese, red: oxygen, purple: potassium.

Recipe for this experiment

Safety, care and cleanliness are important and essential in any chemical experiment. Potassium permanganate ($KMnO_4$) is caustic/irritating to the skin and poisonous if swallowed. Both $KMnO_4$ and H_2O_2 can damage the skin and the eyes. Medi-

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[6] W Potassium_iodide

[7] A detailed explanation is available at W Elephant's_toothpaste. See also Chemiday 3-1-0-124.

[8] Water (H_2O) is a very very stable molecule. It you want to force on more oxygen inside a water molecule, it requires energy to force this union. When H_2O_2 decomposes into H_2O and O_2, this extra energy is what is released as heat. Reaction (1) is an example of an *exothermic reaction*.

cal grade[9] H_2O_2, in fact, is used to disinfect wounds – it really hurts when poured on an open wound! So please be careful when handling these substances. Here's how elephant's toothpaste can be made:

1. Take a clean glass bottle and pour about 30 ml of medical grade H_2O_2. Add few drops of liquid dish soap to this. Keep this bottle in the sink, or in a large-ish dish, because things will spill from the bottle later on. To measure H_2O_2, you can use the measuring cup that comes with many syrupy medicines or a disposable syringe – after discarding its needle safely under adult supervision.
2. Take about 15 ml water in another small container. Add about 2 gm of $KMnO_4$ to this water: This is less than half-a-tea-spoon-full. Dissolve it by stirring it. This solution will have a dark purple color.
3. Pour this solution into the bottle containing H_2O_2 + soap. Elephant's toothpaste will start pouring out of the bottle immediately.
4. Discard[10] these solutions when done, and clean-up!

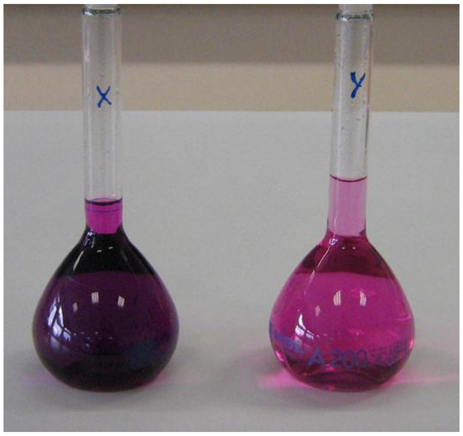

Figure 9: $KMnO_4$ solutions at two different concentrations.

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[9] Typically, 3% or 6%.

[10] For example, by flushing them down the (kitchen) sink. Unfortunately, this is not particularly environmentally friendly, but there is no easy way of safely discarding chemicals like $KMnO_4$!

Figure 10: This beetle, called the Bombardier Beetle, produces H_2O_2 and another chemical called hydroquinone in its abdomen. When threatened, it sprays these two chemicals on its attacker. These react to produce a hot spray with an obnoxious smell.

The gas that comes out of hydrogen peroxide

Earlier, we talked earlier about the two solutions that make elephant's toothpaste. Here, I am going to talk about an idea which I got when I added the second solution ($KMnO_4$ solution) to the first one (H_2O_2 + soap) in a container to make elephant's toothpaste. Without any added color, the foam this produces is white.

Then, I thought: What would happen if I remove the soap from the first solution as well?

In theory, I thought I would get a gas. Looking at reaction (1), that is,

$$2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$$

we might guess that this gas is oxygen. But we never know for sure until we test it ourselves.

So, I took some hydrogen peroxide (H_2O_2) in a test tube and I added some potassium permanganate ($KMnO_4$) to it.

Instantly, a white mist started to rise up from the surface where the reaction was going on. It filled up the test tube completely.

Then I thought of holding a burning matchstick inside the test tube. The moment I did this, the matchstick burnt very brightly and quickly. It did not smell very different than the usual matchstick smoke.

This proves that the gas produced by the reaction must be oxygen. Why? When we burn something in air – for example,

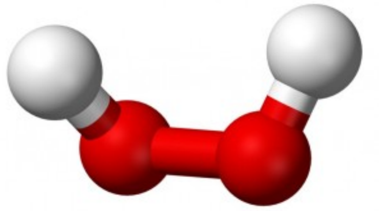

Figure 11: A ball-and-stick model of the H_2O_2 molecule.

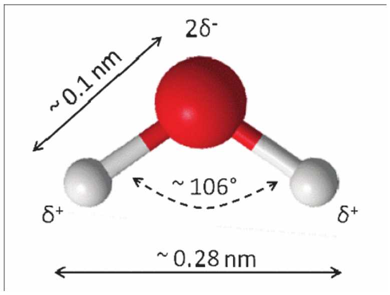

Figure 12: A ball-and-stick model of the H_2O molecule. A nanometer (nm) is 0.00000001 centimeter, that is, 1/100000000-th of a centimeter.

a matchstick – it needs oxygen to burn. But, even though air has about 21% oxygen, I am told that the oxygen available for burning at the base of the flame is only about 5%. And even then, the burning thing has to struggle to get it! So, if the gas produced by the above reaction is indeed oxygen, then the matchstick should burn a lot faster – hence brighter, and without leaving much ash. This proves that the gas produced is indeed oxygen.

Recipe for this experiment

This experiment is basically elephant's toothpaste without the soap. The rising whitish mist is probably best seen in a narrow container like a test tube, but any clear-glass bottle will do. Keep it on a level surface, pour some H_2O_2 in it, and let it settle down. Add a small amount of $KMnO_4$ to it without disturbing the bottle. To verify that the gas produced is indeed oxygen, light up a matchstick and hold it at the mouth of the test tube or the bottle. You need to be careful when playing with fire! This is best done under adult supervision.

Figure 13: Burning splint inside a bottle with concentrated oxygen.

$KMnO_4$: reactant or catalyst?

Then I thought: Does $KMnO_4$ react with H_2O_2 or does it act as a catalyst in this reaction? The reason why I noticed this is because if I put just a tiny amount of $KMnO_4$ in H_2O_2, then its color changes. This could mean that $KMnO_4$ might be changing into something else.

Well, it was difficult to confirm this myself, so I looked up the reaction[1] and found that:

$$2\,KMnO_4 + 3\,H_2O_2 \longrightarrow 3\,O_2 + 2\,MnO_2 + 2\,KOH + 2\,H_2O. \quad (2)$$

Indeed, $KMnO_4$ does react with H_2O_2, and the color change

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[1] Chemiday 3-1-0-356. See also `https://chemistry.stackexchange.com/questions/43256/`.

I noticed is because $KMnO_4$ (deep purple) decomposes into manganese dioxide[2] (MnO_2, brownish). Ideally, one should also check with a litmus paper[3] if the solution is basic[4] because of the potassium hydroxide[5] (KOH) formed as a result of the reaction.

Combustion: What happens when a matchstick burns?

By the way, what kind of chemical reaction do we expect when a matchstick burns? Let's say that a matchstick is mostly wood – except for the coating on its tip. Wood is a kind of sugar called glucose[6] – but not exactly: Wood is mostly made of cellulose[7], cellulose is made of glucose molecules bonded to themselves in a chain called a polymer[8]. Glucose is what keeps us going – it is the source of energy by which our bodies function. So, let's look at what happens when glucose burns:

$$\underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Oxygen}}{6\,O_2} \longrightarrow \underset{\substack{\text{Carbon}\\ \text{Dioxide}}}{6\,CO_2} + \underset{\text{Water}}{6\,H_2O} + \text{Heat} \tag{3}$$

That is, when glucose burns completely, it produces carbon dioxide[9] (CO_2), water (H_2O), and a lot of heat.

So, if some extra concentrated oxygen is available somehow, then this reaction should happen a lot faster than in plain air and produce a much brighter flame.

[2] W Manganese_dioxide

[3] W Litmus_paper. Some internet resources suggest that home-made acidity/basicity indicators can be made from beetroot and turmeric, but I have not tried these myself!

[4] W Base_(chemistry)

[5] W Potassium_hydroxide

[6] W Glucose

[7] W Cellulose

[8] W Polymer

[9] W Carbon_dioxide

Figure 14: Glucose burning in a jar of oxygen.

What is a chemical reaction?

Matter[1], under normal temperatures and pressures, is made of *atoms*[2] and their combinations called *molecules*[3]. An atom has positively charged protons[4] and charge-neutral neutrons[5] in its nucleus[6], and negatively-charged electrons[7] which occupy different *shells* around the nucleus. Molecules are aggregates of atoms which are held together via atomic forces. We will learn a little more about this in the last chapter. These different combinations of atoms and molecules are called *chemical substances*[8].

A *chemical reaction*[9] is a process which transforms one set of chemical substances into another. Chemistry[10] is a grand orchestra of the outer electrons and energy flows[11] through which makes atoms do things. The substances that react in a chemical reaction are called *reactants*, and what they change into are called *products*.

A chemical reaction is shown with an equation like

$$2\,\mathrm{H_2} + \mathrm{O_2} \longrightarrow 2\,\mathrm{H_2O}. \tag{4}$$

We can read this chemical reaction in plain English as

Two hydrogen molecules ($2\,\mathrm{H_2}$)

and ($+$)

one oxygen molecule ($\mathrm{O_2}$)

react to produce ($\longrightarrow$)

two water molecules ($2\,\mathrm{H_2O}$).

"If you don't know chemistry
Here's some on reactions
It may not be much
But it's still a good fraction
I don't know where to start
So I'll begin with this
An element + element yield compound is synthesis
Decomposition is multiple producs from a single reactant"

`https://haikudeck.com/p/ee67a6236e`

[1] W Matter

[2] W Atom

[3] W Molecule

[4] W Proton

[5] W Neutron

[6] W Atomic_nucleus

[7] W Electron

[8] W Chemical_substance

[9] W Chemical_reaction

[10] W Chemistry

[11] W Energetics

By and large, like most natural processes, chemical reactions happen in a direction of greater stability[12] – That is, they reach a point after which nothing changes much. A chemical reaction neither destroys nor creates atoms[13] or energy[14].

Balancing a chemical reaction

Accounting for energy in a chemical reaction is not that easy, but balancing mass in a reaction is easier than you think. How to check that the mass in a reaction is correct? Think of the reaction like you're doing math. For example,

$$\begin{aligned} 1+1 &= 2 \\ \therefore \quad 2 &= 1+1 \end{aligned}$$

The value on the left and the right are the same but, visually, the numbers 1 and 2 look different. But two ones make a two and a two is two ones. The same applies to a chemical reaction. If you start with one carbon[15] atom and two oxygen atoms, you end up with the same number of atoms after the reaction, except that the three atoms form one molecule of carbon dioxide[16]:

$$C + O_2 \longrightarrow CO_2. \tag{5}$$

So, to balance a reaction, you need to count individual atoms on the left-hand side of the " $\longrightarrow$ ", count individual atoms on the right-hand side of the " $\longrightarrow$ ", and see if these numbers match. If they do, then the mass in the reaction is correctly balanced. In the above reaction, we have one carbon atom and 2 oxygen atoms on both the left and right of " $\longrightarrow$ ", so the above reaction is balanced[17].

Let us take a slightly complicated example. Let us look at

[12] W Chemical_stability

[13] W Conservation_of_mass

[14] W Conservation_of_energy

[15] W Carbon

[16] W Carbon_dioxide

[17]

	Left side	Right side
Carbon	1	1
Oxygen	2	2

the reaction when glucose burns (Reaction (3)):

$$\begin{array}{ccccccc} C_6H_{12}O_6 & + & 6\,O_2 & \longrightarrow & 6\,CO_2 & + & 6\,H_2O \\ \text{Glucose} & & \text{Oxygen} & & \text{Carbon} & & \text{Water} \\ & & & & \text{Dioxide} & & \end{array}$$

Here, one molecule of glucose reacts with six molecules of oxygen to produce six carbon dioxide molecules and six water molecules. The glucose molecule has 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen molecules (12 oxygen atoms). The numbers of atoms on the two sides of the " $\longrightarrow$ " are given in the table in the margin[18]: We see that this reaction is balanced. An unbalanced reaction means that some of the atoms involved are not accounted for.

The reason why the numbers of atoms on the two sides are the same is because the reaction does not destroy, fuse, or split the atoms. It only destroys and recreates the chemical bonds[19] that hold the atoms together in the form of molecules.

What is oxidation?

An oxidizer is a chemical substance which oxidizes other substances. Oxidation[20] happens when a chemical substance looses its electrons to another substance. We will try to understand this better in a later chapter. Oxygen happens to be a strong oxidizer. Both the reactions 3 and 4 above are examples of oxidation reactions. In reaction (4), it is hydrogen which gets oxidized. In reaction (3), it is glucose which gets oxidized.

Here's one more example of an oxidation reaction:

$$\begin{array}{ccccc} 2\,Fe & + & O_2 & \longrightarrow & 2\,FeO \\ \text{Iron} & & \text{Oxygen} & & \text{Ferrous Oxide} \end{array} \tag{6}$$

In this reaction, each oxygen atom steals 2 electrons from each iron atom. This is how the iron atom gets oxidized into ferrous

Figure 15: Fast oxidation of a car: Car + Oxygen $\longrightarrow$ Smoke + Flames + Noise + Argument.

---

[18]

	Left side	Right side
Carbon	6	6
Hydrogen	12	12
Oxygen	18	18

.

[19] W Chemical_bond

[20] W Redox

oxide[21]. Iron[22] has a complex chemistry. It can form oxides[23] in many different ways. The rust[24] that we see on iron objects is a mixture of iron oxides and hydroxides.

More generally, oxidation is the taking away of electrons. That can happen even when oxygen is not involved in a chemical reaction. For example, chlorine[25] is another strong oxidizer because it likes to steal electrons from other atoms. The reverse of oxidation – that is, giving away electrons – is called *reduction*[26].

The photograph below (Fig. 16) shows rusty old ships. The salt[27] in the sea water[28] makes iron rust faster.

Figure 16: Rusty old ships in Goa. Salt in the sea water makes iron rust faster.

[21] W Iron(II)_oxide
[22] W Iron
[23] W Oxide
[24] W Rust
[25] W Chlorine
[26] W Reduction_(chemistry)
[27] W Salt
[28] W Seawater

The purple flame

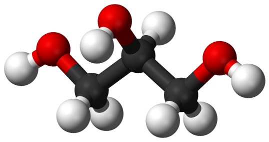

Figure 17: Ball-and-stick model of a glycerol molecule; black: carbon, red: oxygen, white: hydrogen.

Potassium permanganate ($KMnO_4$) is a dark purple, almost-black-looking substance. It is also an oxidizer. Glycerine or glycerol (W Glycerol; see Fig. 17) is a colorless, non-toxic, sweet-tasting, viscous liquid. The glycerine I got has a sweetish odour to it – this was over-the-counter medical-grade glycerine from a medical shop.

The reaction of potassium permanganate with glycerine[1] is an exothermic reaction[2]. An exothermic reaction releases heat. The opposite of a exothermic reaction is an endothermic reaction[3]. An endothermic reaction absorbs heat from the surroundings, and the substance gets cold instead of getting hot. When carrying out such reactions, particularly exothermic ones, we need to be very careful[4].

Because glycerine is a thick, viscous liquid, it does not react with potassium permanganate easily. So, adding a little bit of water helps kickstart the reaction and make it go faster.

When you add water to the potassium permanganate and glycerine mixture, smoke[5] starts to come out after a few seconds. You can also hear a sizzling sound. Smoke becomes thicker and, suddenly, the whole thing lights up into a purplish flame. If you do this experiment yourself, then be careful and keep your hands clean of glycerine and $KMnO_4$.

Apart from the heat generated, this is the reaction that happens:

[1] W Glycerol_and_potassium_permanganate

[2] W Exothermic_reaction

[3] W Endothermic_reaction

[4] See precautions at the end of this chapter.

[5] W Smoke

$$\begin{array}{lllllllllll} 14\,\mathrm{KMnO_4} & + & 4\,\mathrm{C_3H_5(OH)_3} & \longrightarrow & 7\,\mathrm{K_2CO_3} & + & 7\,\mathrm{Mn_2O_3} & + & 5\,\mathrm{CO_2} & + & 16\,\mathrm{H_2O} \\ \text{Potassium} & & \text{Glycerine} & & \text{Potassium} & & \text{Manganese} & & \text{Carbon} & & \text{Water} \\ \text{Permanganate} & & & & \text{Carbonate} & & \text{Trioxide} & & \text{Dioxide} & & \end{array}$$

Recipe for this experiment

Safety, care and cleanliness are important and essential in any chemical experiment. Glycerine by itself is not at all harmful to the skin or even if ingested in small amounts. But when mixed with $KMnO_4$, it will burn the skin. $KMnO_4$, on the other hand, is caustic to the skin and poisonous if swallowed. So do take care when handling these substances!

To perform this experiment at home – under adult supervision – you need

1. potassium permanganate from medical shop, about half a table spoon;
2. glycerine from medical shop, about 10 drops;
3. plain tap water, 2-3 drops;
4. a small china dish or bowl, or a plastic lid which you don't need, or a "dish" made from aluminium foil in the kitchen.

Make sure your hands are clean and dry. Use a spoon to put the potassium permanganate in a small pile in the dish, lid, or the "dish". Keep this in a well-ventilated open place, preferably outdoors. Put glycerine drops on the potassium permanganate pile. Put water drops on top of this. You can use two disposable syringes[6] for glycerine and water. Make sure to move 3 to 5 feet away from this setup. Add a few more drops of water if the reaction does not start in about 10 seconds. The reaction produces some amount of smoke; do not breath it in. After the

[6] Take help from an adult at home and make sure to safely discard their needles first!

reaction is over, the whole thing is very very hot. If you used a plastic lid, it will have melted or even burnt with an obnoxious smell. If you used an aluminium foil “dish”, it will be thoroughly corroded. Wait sufficiently long for to everything cool first, and then safely discard the aluminium foil, the residue in the china dish, etc.

Figure 18: Potassium permanganate + glycerol fire.

The substance that helps things burn

Potassium nitrate (KNO_3)[1] is a strong oxidizer, because it can provide oxygen in a chemical reaction. In powder form, it looks and feels like table salt, but it is a bit whiter than salt. It dissolves in water easily, so making a solution is easy: Just add some potassium nitrate to the water and stir the solution for some time. You might notice that the solution becomes cooler, that is, this particular dissolution process is endothermic.

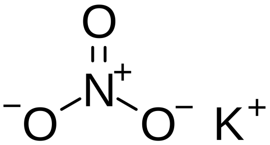

Figure 19: Structure of the potassium nitrate (KNO_3) molecule.

Paper and cardboard are made of wood, which is mostly cellulose, which is in turn a polymer made of glucose monomers[2]. Cotton is also mostly cellulose fibre. So, chemically speaking, paper, wood, cotton and glucose are *nearly* the same when it comes to burning!

Figure 20: Embers on top of burning incense sticks.

Suppose you paint a piece of paper, or a cardboard, or a ball of cotton with potassium nitrate solution. For example, on paper, you could write your name, draw a smiley, or a picture. The solution carries potassium nitrate molecules into the paper or the cardboard or the cotton ball. When the water dries out, potassium nitrate gets stuck there.

Now you can light-up this paper with a glowing/burning ember[3]: Just touch an ember (see Fig. 20) lightly to a part of the paper that has potassium nitrate on it. When lit this way, the paper burns in tiny sparks only where there is potassium nitrate – and etches out your hidden drawing in the paper!

You will notice that the parts of the paper which were soaked in potassium nitrate burn quicker. If you have played with fire

[1] W Potassium_nitrate

[2] W Monomer

[3] W Ember

crackers[4], you might notice that even with potassium nitrate the paper burns much slower than the black powder used in fire crackers. In comparison, a paper without potassium nitrate burns only where you touch the ember. You might also notice that the ash that is left has a different texture than the ash of paper without potassium nitrate, and the smoke produced smells a bit different.

Anything that can burn can be made to burn faster with potassium nitrate. Be careful ... it can also catch fire very easily ... or you might end up burning your house down!

Figure 21: Potassium nitrite (KNO_2) crystals.

There are two reactions which happen when a piece of paper, a cardboard, or a cotton ball (or string) soaked in potassium nitrate solution burns.

First, the heat from the glowing ember – or the burning paper itself – decomposes potassium nitrate into potassium nitrite[5] and oxygen:

$$\begin{array}{llllll} 2\,KNO_3 & \longrightarrow & 2\,KNO_2 & + & \textcolor{red}{O_2} & \\ \text{Potassium} & & \text{Potassium} & & \text{Oxygen} & (7) \\ \text{Nitrate} & & \text{Nitrite} & & & \end{array}$$

This oxygen released from KNO_3 is now available directly for combustion where it is happening. This oxygen is over and above the available oxygen from the surrounding air.

Second, this released oxygen reacts with the paper, the piece of wood, or the cotton ball. For example, the cellulose in these substances, which is a polymer made of glucose monomers, will burn as

Figure 22: Cane sugar (brown or white) varieties contain sucrose, a sugar that is similar to glucose. When crushed – like when eating candy – sucrose produces tiny flashes of light. This is called *triboluminescence*.

$$\begin{array}{lllllll} C_6H_{12}O_6 & + & \textcolor{red}{6\,O_2} & \longrightarrow & 6\,CO_2 & + & 6\,H_2O \\ \text{Glucose} & & \text{Oxygen} & & \text{Carbon} & & \text{Water} \\ & & & & \text{Dioxide} & & \end{array}$$

[4] Remember adult supervision!

[5] W Potassium_nitrite

Recipe for this experiment

As with any chemistry experiment, handling chemicals carefully, cleanly, and being safe are important and essential. The recipe for this exploration is quite simple. Take 2-3 grams of KNO_3 and dissolve it in about 10 ml of water. This is about half-a-teaspoonful, and should make a fairly concentrated KNO_3 solution[6]. Dip a small cotton ball in this solution, drain out excess solution from it, and let it dry thoroughly. Once dry, light it up with an incense stick or a matchstick with a live ember on it. See how quickly it burns without a flame. You might notice an unusual smell as well. Let the burnt cotton ball cool down – it can remain very very hot inside for quite some time. Or pour tap water on it to extinguish it completely. Discard it safely. If you get more adventurous, paint your name on a piece of thick absorbing paper with a brush dipped in the KNO_3 solution. Let it dry. Give another coat and let it dry. When lit up similarly, this will burn your name in the paper[7]!

[6] At 20°C, which is a typical day-time municipal tap water temperature in Pune, about 30 grams of KNO_3 can be dissolved in about 100 grams of water. More such solubility data can be found at `https://periodic-table-of-elements.org/SOLUBILITY`.

[7] See, for example, ▶ Z_I8UG2Xykk.

Figure 23: KNO_3 is one of the ingredients of the otherwise destructive gunpowder. Here's a painting that has been created using gunpowder by artist Danny Shervin. See the image link at the end of this book.

Incomplete burning

Charred charcoal

What we call charcoal[1] and coal[2] are made – mostly – of carbon. Charcoal is made by heating wood pieces in a container (or an oven) with a hole to let out the escaping gases. Essentially, this happens because the air available in the container for burning is not enough for complete or even partial burning.

Figure 24: Charcoal made from wood pieces.

Now, wood is made mostly of cellulose, which is a long chain (polymer) of glucose molecules (monomers), and glucose is in turn made of carbon, hydrogen and oxygen.

Ideally, carbon, hydrogen and oxygen should break apart cleanly leaving only carbon in the container, and the hydrogen and oxygen should combine with each other to form water. In practice, that does not seem to happen, because the escaping gases are combustible – they burn slowly in a yellow flame[3].

This tells us that these gases contain hydrocarbons[4] which burn to form carbon dioxide and water. May be these gases also contain a little bit carbon monoxide[5] CO as well, which burns in bright blue flame. After a while, the flame dies out and the gases stop coming out. This is when your charcoal is ready!

Recipe for this experiment

This experiment can be done easily at home[6]. Here's the recipe.

---

[1] W Charcoal

[2] W Coal

[3] See this link for more details about blue and yellow flames: https://www.elgas.com.au/blog/1585-why-does-a-gas-flame-burn-blue-lpg-gas-natural-propane-methane.

[4] W Hydrocarbon

[5] W Carbon_monoxide

[6] Strictly under adult supervision! If you are not careful, you can get burns on your hands.

1. To do this, you need an empty aluminium soft-drink can, some kitchen aluminium foil and a heat source such as the kitchen stove. Cut the can[7] into two cylindrical pieces with ordinary pair of scissors. With some careful maneuvering, you can fit the top piece over the bottom piece[8].
2. To make charcoal, you can put pieces of paper, cotton balls, or small pieces of a dry stick, etc., in this charcoal "oven"[9]. Put this can on a stove and heat it on a slow stove flame. After gases start escaping from the hole in the top of the can in sufficient quantity, you can light them up! Make sure to not stuff the can too much: For example, cotton balls can be filled to, say, 3/4th height because cotton is a light, fluffy material. On the other hand, if you put denser wood pieces, you must put just a few small pieces. Otherwise, the escaping gases might form a tall flame. Over time, the flame will get shorter as the escaping gases burn out. After the flame gets extinguished by itself and there is very little gas escaping from the can, stop the heat and cover the can with aluminium foil[10] to prevent air from entering the can[11]. The can should cool down in a few minutes. Remove the foil carefully. Open the can. You will find your charcoal inside. You might notice that your charcoal has about the same shape as the original material which you put in, but its size has shrunk down quite a bit. This is because it has lost almost all of its hydrogen and oxygen content.
3. You can now try to light up your charcoal[12]: See if the charcoals made of paper, burnt matchsticks, cotton balls, sticks, cotton cloth, etc., burn differently.

Heavy hydrocarbons

Here's another – and much simpler – experiment about incomplete burning. Instead of collecting carbon, we will now try to

[7] Be careful; the cut edges are usually sharp!

[8] For example, by crimping/indenting the top piece slightly at its bottom edge. Make sure the two pieces fit snugly. Again, all this should be done under adult supervision and very very carefully, because your fingers may get cut if you are not careful (believe me!).

[9] Never use any plastics or other artificial material: These will give out horrible and possibly toxic gases. They will also ruin your "oven" by leaving a hard residue inside.

[10] Here's a suggestion: Before the experiment, cover the can with crumpled aluminium foil and get the foil into the can's shape, so that when the can is hot, it is easier to put the foil cap on the can.

[11] If air enters the can, the hot charcoal inside will smoulder and might even burn.

[12] Again, under adult supervision!

collect the hydrocarbons! You will need a glass mug or bowl, a piece of newspaper, a fork[13], and a matchbox or a lighter.

1. Tear the newspaper up into narrow strips; say, about 7 cm in width and about 9 cm in length. The exact size does not matter. Roll the strip into a cylindrical roll which is neither too tight nor too loose.
2. Insert the rolled strip between two spikes of the fork with most of the cylinder on one side of the fork.
3. Light up the paper cylinder at the longer end and hold the shorter end inside the glass. You will see some gases escaping upwards from the flame. Interestingly, after a few seconds, a heavier gas[14] starts escaping from the bottom and gets collected in the container below. If you can collect sufficient amount of this gas, it often looks yellowish white. You will notice that it is flammable: It is likely to be a mixture of carbon monoxide (CO) and hydrocarbons which are heavier than air.
4. I suggest playing around with the length and width of the paper strip and the diameter of the cylinder you roll it into. See if you get different results with different sizes.

[13] A fork which you do not need!

[14] It is heavier probably because it is cooler.

Making flammable gel

Now you might be thinking, “How can a jelly burn?” Well, it can; it depends on what it is made-up of. The one that I made had an alcohol in it. This experiment is inspired by two videos[1] I saw on YouTube, and it can be very easily done at home.

This flammable gel is made by mixing solution of calcium acetate[2] ($Ca(CH_3COO)_2$) in water with any alcohol. Calcium acetate is soluble in water and not soluble in alcohol. When these two are mixed, the calcium acetate particles get evenly distributed in the solution and form a kind of a net, which gives it a semi-solid, gel-like form. Water and alcohol molecules are kind of trapped inside this net. This gel is somewhere in between a solid and a liquid.

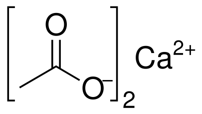

Figure 25: Structure of the calcium acetate ($Ca(C_2H_3O_2)_2$) molecule.

What is a gel[3]? A gel is a substance where two or more substances mix well with each other but do not react with each other or dissolve into one another the way salt dissolves into water. Gels are a kind of *colloids*[4]. Milk is a colloid.

In our case, the calcium acetate molecules form a lightly rigid structure in the container, allowing it to retain its shape. Alcohol, which cannot dissolve calcium acetate and the water that we used to make the calcium acetate solution occupy the empty spaces inside this structure. See Fig. 26.

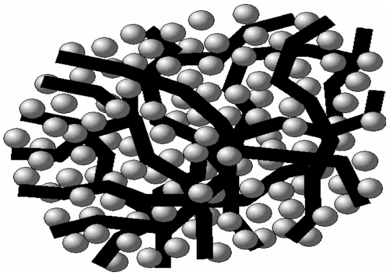

Figure 26: Schematic structure of a gel/colloid.

Recipe for this experiment

1. Get any type of alcohol with 99% concentration[5]. Alcohol

---

[1] ▶ DjmnSoupVoA ▶ 32vCLXTjnyQ

[2] W Calcium_acetate

[3] W Gel

[4] W Colloid

[5] Do NOT use methanol (W Methanol) even if you can get hold of it, because it is very toxic.

diluted with water will not work. I used 99% isopropanol (which is same as isopropyl alcohol). Use a non-toxic alcohol like ethanol or isopropanol[6].

2. How to make calcium acetate? Firstly, we need white vinegar (CH_3COOH). The main chemical ingredient of vinegar is acetic acid. The kitchen variety will do, but better if it is 20% or higher. Secondly, we need a source of calcium carbonate[7]. I used egg shells[8]. After some experimentation, I found that mildly crushed egg shells work better instead of grinding them into a fine powder[9].

   To make calcium acetate, take a container and fill it half way through with vinegar. If you fill it too much, then it will start to spill when you add egg shells (or calcium carbonate in any other form). Put the egg shells in vinegar. In the beginning, some of these egg shells might sink and some might float. Stir the mixture so that all sink.

   You will notice that egg shell pieces start to form little bubbles around them. This is because calcium carbonate reacts[10] with vinegar to form carbon dioxide:

$$CaCO_3 + 2\,CH_3COOH \longrightarrow Ca(CH_3COO)_2 + H_2O + CO_2. \tag{8}$$

   The pieces will start to rise and fall just like a lava lamp. After a while, all the pieces will float back to the top. You will need to stir again. I suggest stirring it a few times every ten minutes or so. To make sure all the vinegar in it is used up, keep it overnight. The next morning, filter it through a piece of tissue paper to make sure that contaminants and remaining egg shells are removed. To make it more concentrated, boil the solution for a little while and let it cool down. You might still find stuff floating on this solution – especially if you used egg shells. Filter it again. Now you should be left

Figure 27: Crystals of a naturally-occurring form of calcium carbonate, known as the mineral calcite.

Figure 28: Marble Rocks, Jabalpur, India. Marble is a metamorphic rock which is mostly calcium carbonate.

---

[6] As medical spirits, these are diluted. Better if you can get it in a concentrated form.

[7] W Calcium_carbonate

[8] You can watch this video: ▶ 32vCLX-TjnyQ

[9] Using a kitchen mixer!

[10] Chemiday 3-1-0-12937

(mostly!) with calcium acetate solution.

3. Now, to make about 100 ml of flammable gel, take about 75 ml of alcohol and about 25 ml of our calcium acetate solution. However, calcium acetate solution, the way have prepared it, may not be concentrated enough. So, if you see any alcohol remaining at the bottom, simply add a little more of the calcium acetate solution to the alcohol directly. Now, if calcium acetate solution remains, then add some alcohol to it. All this is trial and error. Any excess alcohol will evaporate over time, so I suggest keeping it in a container in a cool and dry place or in a refrigerator[11]. You should see a gel-like substance formed. This is our flammable gel, finally!

You can take a small chunk of it (say, a spoonful), put it in a dish and light it up with a burning matchstick. You will notice that the flame keeps on burning for a long time. As it burns, some white stuff starts to appear on the surface of the gel. That white stuff, I think, is calcium carbonate. After a while, when the white stuff starts covering the surface of the gel, the flame starts to die out. This is because this white layer starts blocking the alcohol or oxygen supply. Use any metal object – a kitchen knife or a spoon – to cut the gel piece into smaller pieces. This will now expose fresh surfaces for the alcohol to escape (or for oxygen to reach), and the flame will immediately grow bigger.

As with all such experiments: Safety first, and adult supervision a must!

[11] Here's my observation: If you keep the solution for too long, it turns blackish. I suspect, but I am not sure, that this is might be fungus growing around organic impurities, tiny egg shell pieces, etc., despite all the alcohol in the gel!

The physics of chemistry

Chemistry is one hell of a game that atoms and molecules like to play with each other using their outer electrons[1]. How do we try to understand it?

One way is to explore, experiment, record observations, try to guess the patterns in the observations, and try to find explanations for those patterns based on what is already known. This is how our science has evolved, historically.

Another – quicker – way is to start from what we already know about the physics of atoms and molecules, and try to find explanations from there. Let us take the second short-cut approach here.

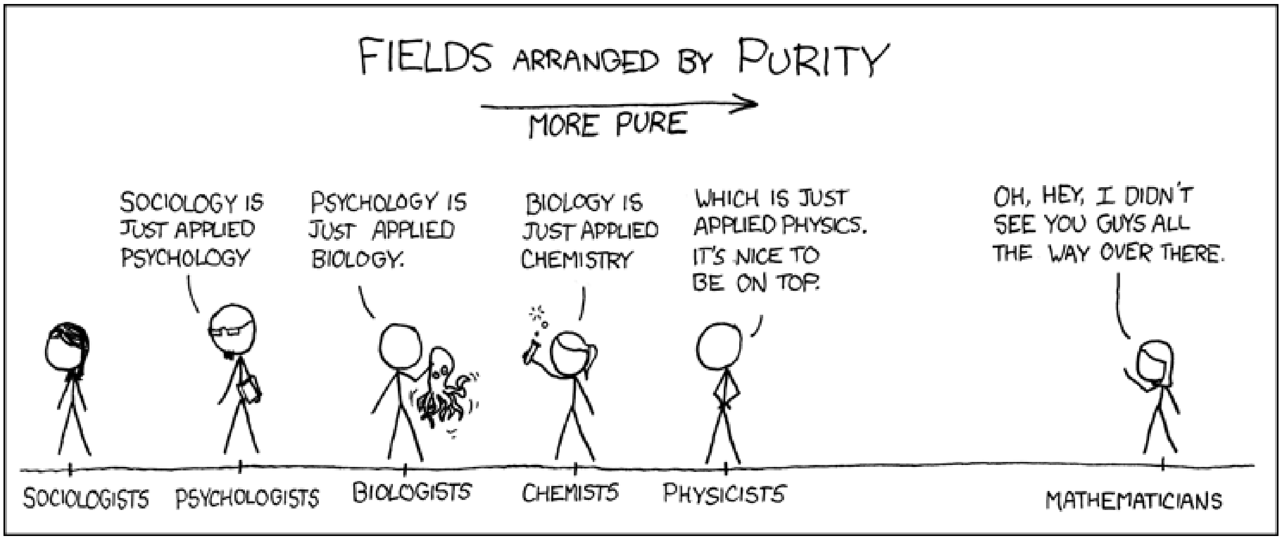

Figure 29: XKCD's comical view of the perceived relationship between natural sciences, humanities and mathematics.

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[1] Other outer-electron games include electric conduction, magnetism, superconductivity, alloying, optics, mechanics, friction, etc.

What is ordinary matter made up of?

Ordinary matter[2], under ordinary conditions, is made up of atoms[3] and molecules[4]. Molecules are aggregates of atoms sticking together. Atoms are made of electrons[5], protons[6], and neutrons[7]. An electron has a negative electrical charge[8]. A proton has an equal but opposite – that it, positive – electrical charge. To make the atom stable – and hence neutral – it has to have as many electrons as there are protons, so that the net electrical charge on an atom is zero. Protons and neutrons sit in the nucleus[9]. A neutron has no electrical charge, so the net electrical charge does not change with any number of neutrons. A neutron has nearly the same mass as a proton. The mass of an electron is about 1/2000th of the mass of a proton. Atoms of different chemical elements[10] have different numbers of electrons (and protons). For example, an iron atom has 26 electrons and 26 protons, and an oxygen atom has 8 electrons and 8 protons. The number of electrons (or protons) in an atom is called its *atomic number*[11]. The same chemical element is sometimes available in different varieties called *isotopes*[12]. These different varieties have the same atomic number, but they have different numbers of neutrons. For example, iron comes in 4 different stable varieties[13] with the same atomic number 26 which have 22, 24, 25, and 26 neutrons each. Protons and neutrons huddle together in a tiny ball – the nucleus – at the center of an atom.

Hotel Aufbau

Electrons swarm around the nucleus in different shells and orbitals following the laws of quantum mechanics[14]. How are these shells and orbitals arranged? They are arranged according to their *energies*[15]: It takes certain amount of energy to be in an orbital. This is like the rent that has to be paid to live in a

[2] W Matter

[3] W Atom

[4] W Molecule

[5] W Electron

[6] W Proton

[7] W Neutron

[8] W Electrical_charge

[9] W Atomic_nucleus

[10] W Chemical_element

[11] W Atomic_number

[12] W Isotope

[13] The word *stable* here means that these four varieties are not radioactive (W Radioactive_decay).

[14] W Quantum_mechanics

[15] W Energy

hotel room. How do electrons get filled up in these shells and orbitals? By energy, that is, by rent. To make the atom as stable[16] as possible, electrons need to be filled up from the bottom floor to the top floor in Hotel Aufbau. This is called the *Aufbau principle*[17]. It is illustrated in Fig. 30.

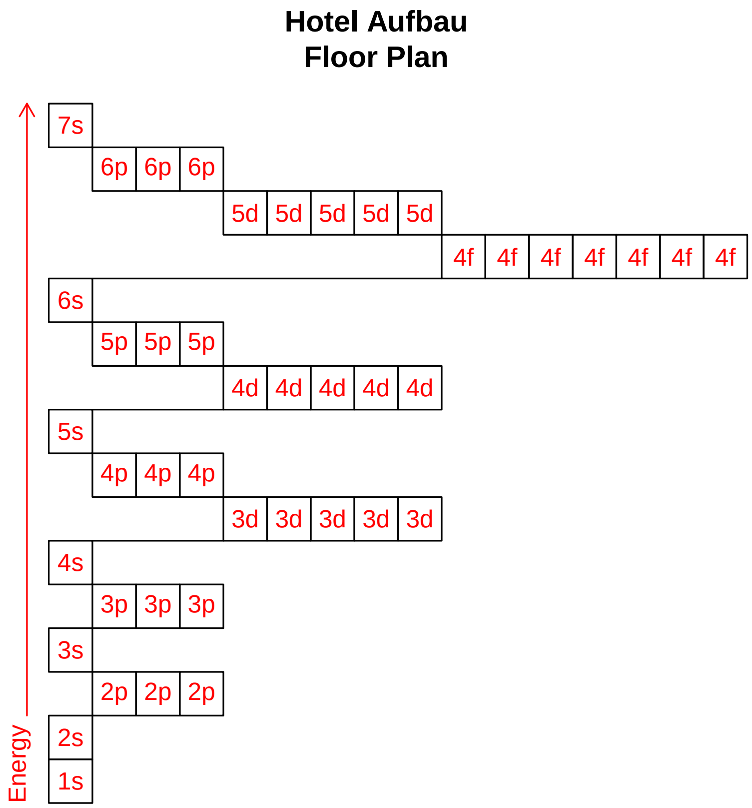

Figure 30: The Aufbau Hotel. The horizontal axis does not mean anything: It is used only to spread the s, p, d, f orbitals out in the schematic plan of the hotel.

The first shell, known as K or 1, has a single orbital called s. The s orbital can hold up to 2 electrons[18]. The second shell, known as L or 2, has two orbitals called s and p. The p orbital can hold upto 6 electrons in its 3 suborbitals. The third shell, known as M or 3, has 3 orbitals called s, p, and d. The d orbital can hold upto 10 electrons in its 5 suborbitals. The fourth shell, known as N or 4, has four orbitals called s, p, d, and f. The f orbital can hold upto 14 electrons in its 7 suborbitals[19]. This is how it continues to higher shells or floors of Hotel Aufbau: For

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[16] This refers to what is called the quantum mechanical ground state (W Ground_state) of the atom.

[17] W Aufbau_principle. *Aufbau* is a German word which means construction or build-up.

[18] If there are two electrons in an s orbital, they have to have antiparallel spins. Spin is a quantum mechanical property of particles like electrons, protons, neutrons, etc.

[19] Each suborbital of $p, d, f, \ldots$ can hold up to 2 electrons with antiparallel spins.

example, the fifth shell has 9 orbitals, and so on and so forth.

Using the Aufbau principle, the electronic configuration of the neutral iron atom (Fe) is shown in Table 1. Similarly, the electronic configuration of the neutral oxygen atom (O) is shown in Table 2.

$_{26}$Fe	s	p	d	f
1 (K)	2	—	—	—
2 (L)	2	6	—	—
3 (M)	2	6	6	—
4 (N)	2	0	0	0

Table 1: Electron fill-up for $_{26}$Fe. This is also written as $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^6\,4s^2$.

$_{8}$O	s	p
1 (K)	2	—
2 (L)	2	4

Table 2: Electron fill-up for $_{8}$O. This is also written as $1s^2\,2s^2\,2p^4$.

These are called the *ground-state configurations*. They have the smallest possible energy. Altered configurations, where an electron is moved from a filled orbital in the ground-state configurations to an empty orbital are called excited states. These have higher energies than the ground-state configurations.

The stability of different atoms can be different. For example, the atoms (Table 3) are particularly stable because they have completely filled shells. Because of this, these atoms rarely get involved in chemical reactions. So, they are called the *noble* or *inert* elements, atoms or gases[20].

Atom	Electronic configuration
$_{2}$He	$1s^2$
$_{10}$Ne	[He] $2s^2 2p^6$
$_{18}$Ar	[Ne] $3s^2 3p^6$
$_{18}$Kr	[Ar] $3d^{10}\,4s^2 4p^6$
$_{54}$Xe	[Kr] $4d^{10}\,5s^2 5p^6$
$_{86}$Ra	[Xe] $4f^{14}\,5d^{10}\,6s^2 5p^6$

Table 3: Electronic configurations of the noble gas atoms. Here, [He], [Ne], ... mean the electronic configurations of helium, neon, ...

From atoms to molecules

Sometimes aggregates of more than one atom – of the same kind or of different kinds – can be more stable than single atoms. When atoms stick together, their aggregates are called molecules.

Let's talk about table salt[21], which is (mostly) sodium chloride (NaCl). You can see the electronic configurations of the sodium[22] and chlorine[23] atoms in Table 4.

Atom	Electronic configuration
$_{11}$Na	[Ne] $3s^1$
$_{17}$Cl	[Ne] $3s^2 3p^5$

Table 4: Electronic configurations of Na and Cl atoms. Chlorine can be seen as one electron short of the stable neon configuration; sodium can be seen as one electron extra over the stable neon configuration.

These configurations are written with reference to the previous noble gas atom, which happens to be neon[24] for both. We see that sodium has one extra electron above neon, and chlorine has one less electron than neon. This makes both these atoms particularly vulnerable and hence reactive. Sodium prefers to give away its extra electron and slide down into the closed-shell configuration of neon. Chlorine, on the other hand, prefers to snatch an extra electron and slide up into the closed-

---

[20] W Noble_gas

[21] W Salt

[22] W Sodium

[23] W Chlorine

[24] W Neon

shell configuration of argon (see Table 3). So, when sodium and chlorine atoms come together, they like to exchange electrons, get charged up, and stick to each other to form NaCl!

Another example. The last orbital (p) oxygen has 4 electrons but it has the capacity to hold six and filling that orbital gives oxygen extra stability. So oxygen likes to grab electrons. On the other hand, iron $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^6\,4s^2$ electronic configuration[25]. But iron has two electrons in its outermost ($4s$) shell so it would be more stable if if didn't have any electrons in that shell at all. So it wants to give away those two electrons. So, when an iron atom and an oxygen atom come together, they like to stick together because iron gives away two electrons and oxygen likes to have two extra.

This way the NaCl molecule is held together is called *ionic bonding*[26]. This is one kind of arrangement for sharing electrons. The electron-sharing arrangements between different atoms can be quite complicated[27]. For example, carbon engages with other atoms using what is called *covalent bonding*[28].

What is chemistry?

When atoms and molecules meet in a shared environment, they can negotiate more stable or less stable rearrangements. This can produce a different sets of atoms and molecules through rearrangements of the atoms involved. We call this a *chemical reaction*. Chemistry is the science which tries to make sense of how atoms and molecules interact with one another through chemical reactions!

[25] Iron has 2 electrons in $4s$ instead of $3d$ because $4s$ has less energy than $3d$ does with 10 electrons.

[26] W Ionic_bonding

[27] W Chemical_bond

[28] W Covalent_bond

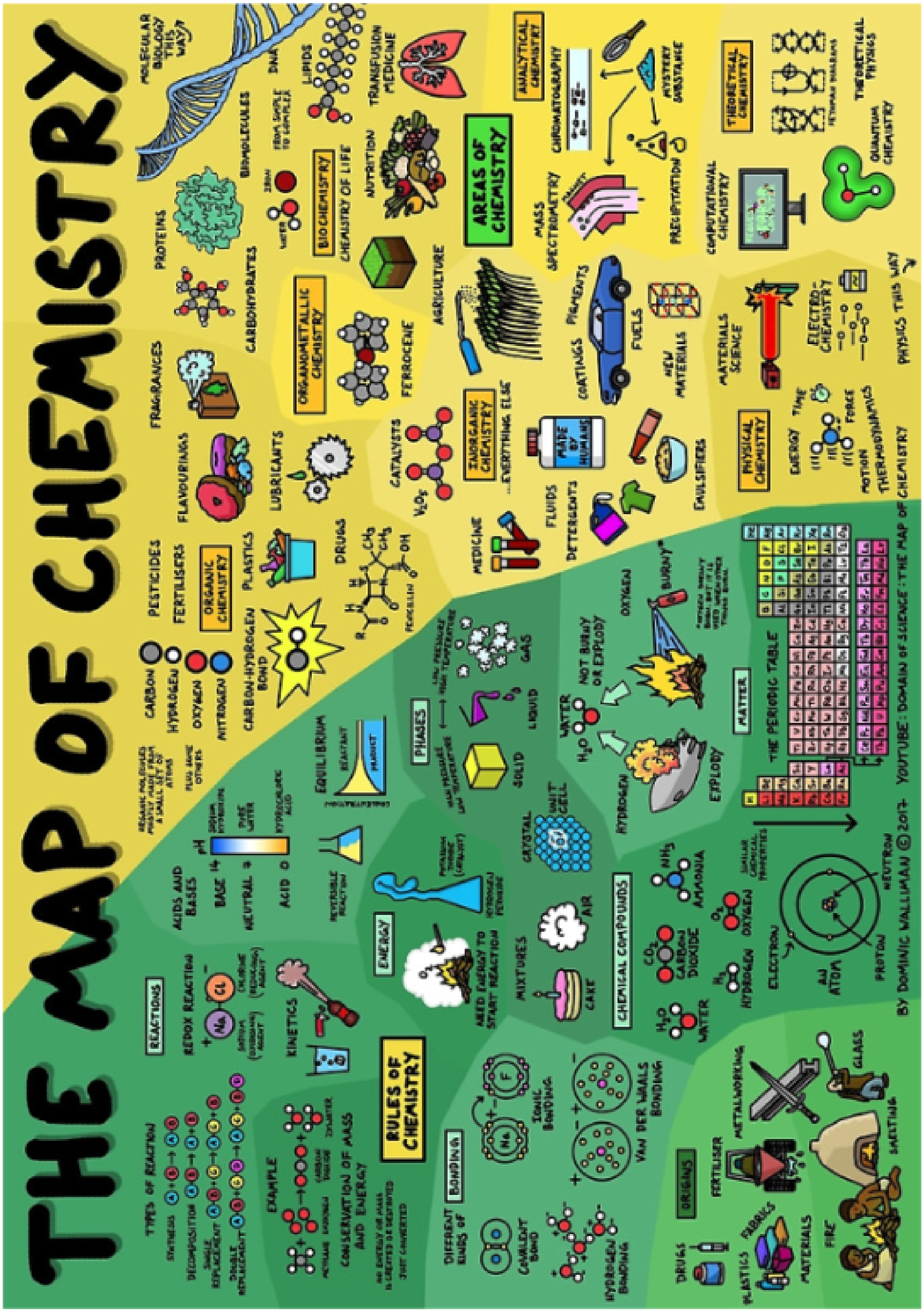

Figure 31: Dominic Walliman's Map of Chemistry

Image credits

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FIGURE	URL
Cover	https://www.ebay.co.uk/itm/A3-Awesome-Chemistry-Poster-Size-A3-Chemist-Science-Art-Poster-Gift-14791-/163736021175
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7	https://en.crystalls.info/Potassium_iodide
8	https://en.wikipedia.org/wiki/Potassium_permanganate
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NaCO2
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SCIENCE