Appendix 2 - Background Contents

1. Energy and power

Energy

This module is about energy. Energy is needed to move an object when a force opposes that movement, to heat something, or to emit light (radiation).

Power

When talking about energy, the concept of power is often important

Power is the amount of energy converted per second.

In formula:

Formula (1) $\quad P = \dfrac{(\Delta)E}{t} \Rightarrow E = P \cdot t$

In which:

- P = power in watts (W)
- E = energy in Joules (J)
- t = time in seconds (s).

Power plants often differentiate between average power and peak power. Wp is then used as the unit for peak power.

Units of energie

The SI unit of energy is Joules. Joule is also called the watt second (Ws).

- 1 joule is the amount of energy required to lift an object weighing 100 grams 1 meter.
- 1 joule (of Ws) is the energy converted when a device with an output of 1 watt is turned on for 1 second.

This shows that a joule is very little. That is why the unit kWh (pronounced kilowatt hour) is usually used for electrical energy.

- 1 kWh is the amount of energy converted when a device with a power of 1000 watts (one kilowatt) is turned on for 1 hour.

So there are two ways, in which you can calculate with the relationship between energy and power ($E = P \cdot t$):

- $E\left(\text{in J}\right) = P\left(\text{in W}\right) \cdot t\left(\text{in s}\right)$

- $E\left(\text{in kWh}\right) = P\left(\text{in kW}\right) \cdot t\left(\text{in h}\right)$

To calculate how many joules a kilowatt hour is, we also use the relationship between energy and power and the fact that 1 hour lasts 3600 seconds:

$$E = P \cdot t \Rightarrow 1\ \text{kWh} = 1000\ \text{W} \ \cdot \ 3600\ \text{s} = 3600000\ \text{J}\ .$$

Prefixes E (in J)

Because power generation involves large amounts of energy or power, larger prefixes are often used:

Tera (T)	= triljon	= 1 000 000 000 000	= 10^{12}
Giga (G)	= billion	= 1 000 000 000	= 10^{9}
Mega (M)	= million	= 1 000 000	= 10^{6}
Kilo (k)	= thousand	= 1 000	= 10^{3}

2. Energy and labor

To move an object or to keep it moving (on Earth), energy is always needed. Therefore, forces on a moving object always cause an energy conversion (at least, as long as the force is not perpendicular to the direction of motion). We say physically: a force on a moving object produces work. The work delivered is a measure of the amount of energy converted during motion.

The amount of labor produced depends on the magnitude of the force and the path traveled (provided the direction of F and s are parallel):

Formula (3) $\quad W = F \cdot s$

Where:
- W = labor in joules (J)
- F = force in newton (N)
- s = distance traveled in meters (m).

For example, for a falling object, the energy converted (ΔE) during the fall over a distance (Δh) is equal to:

$$\Delta E = W = F_z \cdot s = m \cdot g \cdot \Delta h$$

3 The mole, the unit of chemical quantity

There are several ways to indicate quantities of substance, such as units of mass (ton, kg, g, mg) and units of volume or content (m3, dm3, cm3, L, mL). Chemists often use the unit mole. To understand this mole, knowledge of atomic masses (formerly: atomic weights) and molecular masses is necessary. The table in Figure 10 gives important atomic and molecular masses, expressed in the atomic mass unit u. This is the mass of the lightest atom, the hydrogen atom:

Atoom	Atoommassa (in u)
H	1
C	12
O	16
Molecuul	Molecuulmassa (in u)
H_2	2
O_2	32
H_2O	18
CH_4	16
CO_2	44

Figure 10: atomic and molecular masses.

By definition, a mole is the number of atoms in 12,000 g of 12C, so that is the atomic mass of carbon, but now expressed in grams. This is obviously an incredibly large number, very roughly rounded off: 600 000 000 000 000 000 000 000. This 6 with 23 zeros, more conveniently written as 6×1023, is called Avogadro's constant (the number).

Thus, according to the definition, you have a mole of substance if you put the unit 'gram' after the molecular mass of that substance. From the atomic masses H = 1 and O = 16 follow the

molecular masses and thus the molar masses (also called: molar masses) of hydrogen (H_2), oxygen (O_2) and water (H_2O). These are 2 g, 32 g and 18 g, respectively.

The great advantage of the mole is that it bridges the gap between the atomic scale and the measurable scale. A molecule of water is not measurable, but a mole of water is: 18 g.

With gases, there is something special going on. The volume of a mole of gas, the molar gas volume V_{molar} is constant for all gases at a certain temperature and pressure. At 20 oC and standard pressure, this constant is 24 dm^3 mol^{-1}.

The mole is also used to indicate a concentration (= content) of particles in a solution. For solids in liquids, a combination of mass and volume is obvious. For example, you can specify the table salt concentration in seawater as 23 g L^{-1}. The most commonly used unit of chemical concentration in solutions is the mole per liter or molar (mol L^{-1}, abbreviated M). Beware: this molar is not the same as the molar in molar mass!

In chemistry, one often indicates the concentration by the formula of the solute in square brackets. For example, for an ammonia solution this becomes: $[NH_3]$ = 0.20 mol L^{-1} = 0.20 M. Those square brackets give real particles present. For example, for table salt, if you have dissolved 0.15 mol of NaCl per liter, this means: $[Na^+Cl^-]$ = 0 M (since the salt is split into ions on dissolution), $[Na^+]$ = $[Cl^-]$ = 0.15 M.
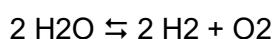
4 Chemical Equilibrium

Batteries provide power, but you can also recharge many batteries. There are sunglasses whose lenses turn dark in the light, light in the dark. You can decompose water into hydrogen and oxygen, but you can also make water from hydrogen and oxygen. The similarity in these very different situations is that the reactions that occur can go in either direction, they are reversible ('reversible'). We take a closer look at the situation with water.

Decomposition: $2 H_2O \rightarrow 2 H_2 + O_2$

Formation: $2 H_2 + O_2 \rightarrow 2 H_2O$

At certain temperatures (in this case above 1500 oC) the two reactions occur simultaneously:

$$2 H_2O \leftrightharpoons 2 H_2 + O_2$$

Appendix 2 - Background Contents

One speaks of chemical equilibrium here; the double arrow is pronounced "in equilibrium with." It is also a dynamic equilibrium because the two reactions continue after equilibration. However, the concentrations of the reacting substances no longer change; as much water is decomposed as is formed in a given time. The conditions for equilibrium are consistent with the preceding story about energy and entropy. Exothermic reactions can be reversible if entropy decreases, endothermic if entropy increases.

There are homogeneous and heterogeneous equilibria. In **homogeneous equilibria**, all reacting substances are in solution or in the gas phase; in **heterogeneous equilibria** they are not. Then, for example, a solid reacts with a dissolved substance.

Almost all reactions are reversible. That a reaction is nevertheless considered to be decaying (= non-reversible) can have several causes. These are the temperature (as in the example of water), a very one-sided equilibrium position and a substance that withdraws from the equilibrium. With a very one-sided equilibrium position, the amount of substances on one side of the equilibrium sign is negligible. The withdrawal of a substance from the equilibrium mixture is well illustrated in the decomposition of limestone (calcium carbonate), a heterogeneous equilibrium:

$$Ca^{2+}CO_3^{2-}(s) \quad Ca^{2+}O^{2-}(s) + CO_2(g)$$
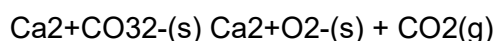
In a closed space, this equilibrium sets when heated. In an open space, the carbon dioxide gas escapes; the equilibrium then tapers off to the right. This is how quicklime (calcium oxide) used to be prepared from shells (main component limestone) in lime kilns.
The redox reactions discussed in this module are also reversible in principle.

5 Logarithms

The logarithm is a mathematical function, usually abbreviated to log. The logarithm of a number is calculated from a base number. There are an infinite number of possible logarithms. Practically, two systems are used. These can also be found on most calculators.

- Logarithms with a base of 10. The Briggs logarithm is used and written as log.
- Logarithms with base e. One speaks of natural logarithm, or Neperian or Neperse logarithm, after the inventor John Napier. The natural logarithm is often noted as ln.

Appendix 2 - Background Contents

The logarithm for the base number a of a number x is the power to which one must raise the base number to get x as the result, i.e. :

$$q = \log_a(x) \Leftrightarrow a^q = x.$$

Or written differently:

$$a^{\log_a(x)} = x.$$

Calculation rules

Already centuries ago, the logarithm was important for people who had to do a lot of arithmetic.

A property of logarithms is that a multiplication can be converted to an addition:

$$\log(a) + \log(b) = \log(ab)$$

Dividing: by subtracting logarithms, divisions can also be performed.

$$\log(a) - \log(b) = \log\left(\frac{a}{b}\right)$$

Power multiplication: if you take the logarithm twice in a row, you can raise a number to an arbitrary power by adding.

$$\log\big[\log(a)\big] + \log(b) = \log\big[\log(a^b)\big]$$

Roots: if the logarithm is taken twice in succession, then any root can be obtained by subtracting.

$$\log\big[\log(a)\big] - \log(b) = \log\left[\log\left(a^{\frac{1}{b}}\right)\right] = \log\big[\log(\sqrt[b]{a})\big]$$

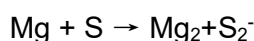### 6 Redox reactions

A common classification of reaction types gives the tripartite acid-base reactions (transfer of protons, H+), the redox reactions (transfer of electrons, e-) and a residual group important for this module.

Originally, we spoke only of oxidation in a reaction with oxygen, such as the combustion of magnesium:

$2\ Mg + O_2 \rightarrow 2\ Mg^2 + O_2^-$
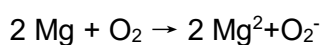
A magnesium atom donates two electrons to an oxygen atom. This also happens in the reaction of magnesium with sulfur, for example:
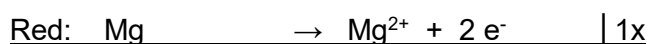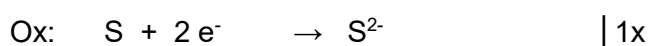
$Mg + S \rightarrow Mg_2 + S_2^-$

So you can broaden the concept of oxidation by linking it not exclusively to oxygen, but to electron transfer. We then speak of redox reactions (redox is the abbreviation for reduction-oxidation), which are often split into two half-reactions:

$Mg \rightarrow Mg_2^+ + 2\ e^-$

$S + 2\ e^- \rightarrow S_2^-$
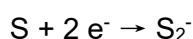
The particle that gives up electrons in a redox reaction is called the reductor, the particle that takes up electrons the oxidizer. You represent the process as follows:
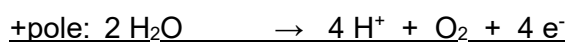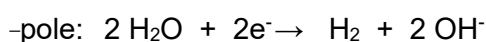
Ox: $S\ +\ 2\ e^- \quad \rightarrow \quad S^{2-}$ | 1x
Red: $Mg \qquad\qquad \rightarrow \quad Mg^{2+}\ +\ 2\ e^-$ | 1x
$Mg\ +\ S \qquad \rightarrow \quad Mg^{2+}S^2$

The two particles involved in a half-reaction are called a redox couple (here they are the couples Mg/Mg_2^+ and S/S_2^-). Binas table 48 lists redox couples with their oxidizer strength. The particles are arranged by decreasing oxidizer strength, top left are the strongest oxidizers and downward the oxidizer strength decreases. On the bottom right are the strongest reductors and towards the top the reductor strength decreases.
A redox reaction proceeds spontaneously when the oxidizer is above the reductor. If the oxidizer and reductor are near each other, an equilibrium reaction occurs. If the oxidizer is well below the reductor, no reaction occurs.
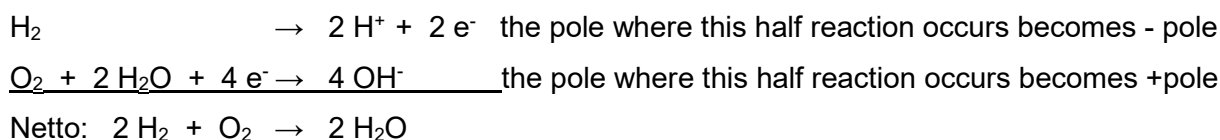
In the above described reactions and also for example in rusting (generally: corroding) the electron transfer takes place directly from one particle to the other. It can also be done differently. In electrolysis and in the reactions in electrochemical cells (in short: cells), each half-reaction takes place at an electrode, i.e. separately from each other. Electron transfer occurs via the connecting wires.
In electrolysis (literally: decomposition by electricity), electrical energy is converted into chemical energy. In other words, by a voltage difference at the electrodes you get chemical reactions. For example, the electrolysis of water:

$-$pole: $2\ H_2O\ +\ 2e^- \rightarrow\ H_2\ +\ 2\ OH^-$
$+$pole: $2\ H_2O \qquad \rightarrow\ 4\ H^+\ +\ O_2\ +\ 4\ e^-$
Netto: $2\ H_2O \qquad \rightarrow\ 2\ H_2\ +\ O_2$

On the contrary, in cells, which are the focus of this module, chemical energy is converted into electrical energy. In other words, a chemical reaction gives you a voltage difference at the electrodes. Take for example a hydrogen cell, where the following half-reactions take place:

$H_2 \qquad\qquad \rightarrow \ 2\,H^+ + \ 2\,e^-$ the pole where this half reaction occurs becomes - pole

$\underline{O_2 \ + \ 2\,H_2O \ + \ 4\,e^- \rightarrow \ 4\,OH^-}$ the pole where this half reaction occurs becomes +pole

Netto: $2\,H_2 \ + \ O_2 \ \rightarrow \ 2\,H_2O$

7 Wire resistance

How great the resistance of a wire is depends on three parameters:

- *Length* (l, in m): the greater the distance electrons travel through the material, the greater the resistance they encounter. Thus, the longer a material is, the greater its resistance. The resistance is proportional to the length: R ~ l.
- *The cross-sectional area* (A, in m2): the greater the area of the (perpendicular) cross-section, the less resistance the electrons experience. The resistance is inversely proportional to the area: R ~ 1/A.
- *Specific resistance* (ρ in Ωm2 m-1 = Ωm): an iron wire and a copper wire of equal length and cross-sectional area differ in resistance. Each material has its specific resistance. The specific resistance of a material is the resistance of a "wire" of that material 1 m in length and 1 m2 in cross section: R ~ ρ.

The specific resistances of metals and alliants depend on temperature and are listed in Binas.

The three factors are combined in Pouillet's law:

$$R = \rho \cdot \frac{\ell}{A}$$

Calculation example

A copper pipe has a cross section of 4 mm^2 and is 200 m long. What is the resistance of that pipe at 20°C?

Answer: given l = 200 m, A = 4 mm2 = 4·10-6 m2. The specific resistance of copper at 293 K is given in Binas: ρ= 17·10-9 Ωm. If we enter these values into Pouillet's law, the resistance follows from:

$$R = \rho \cdot \frac{\ell}{A} = 17 \cdot 10^{-9} \times \frac{200}{4 \cdot 10^{-6}} = \frac{3.4 \cdot 10^{-6}}{4 \cdot 10^{-6}} = 0.9 \ \Omega$$