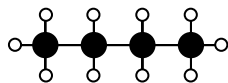


recommended background capabilities

general aptitude: recognising patterns, and interpolation and extrapolation

C1. Stoichiometry

1. The combustion of butane in an oxygen-rich environment produces carbon dioxide and water.



butane



oxygen

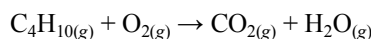


carbon dioxide



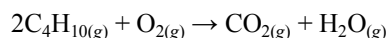
water

The following chemical equation describes the combustion of butane. It is not balanced, because it shows different numbers of each kind of atom on each side of the equation.



Each chemical symbol represents an atom of the corresponding element. Subscript numbers represent how many of that kind of atom are present in a molecule. We combine these symbols and subscripts into chemical formulas, for each of the molecules involved in this chemical reaction.

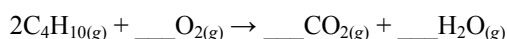
We can write numbers as a prefix to the chemical formulas, to mean that a certain number of molecules are involved in the reaction. This number is called a “coefficient”. For example, we have a coefficient of 2 for butane (C_4H_{10}) below, indicating that two molecules of butane participate in the reaction.



The equation now shows a total of 8 carbon atoms on the left, and 20 hydrogen atoms. To properly balance this equation, we need to match the number of each type of atom on each side of the equation.

#

1. Balance the equation for combustion of butane, by using the correct coefficients for each molecule, to match the number of each type of atom on each side of the equation.



Confidence: 0 | 1 | 2 | 3 | 4

2. Molecules are very small. There are about 20 000 000 000 000 000 000 000 000 (20 million million million million) molecules of oxygen (O_2) in one kilogram of oxygen. These large numbers are inconvenient to work with, so instead of measuring chemicals in “number of molecules”, we use a unit called the “mole”.

The mole unit is convenient because it matches atomic masses (which we can look up on most reference periodic tables) to masses measured in grams. The relative atomic mass of helium is 4.003 atomic mass units (amu), and the mass of 1 mole of helium is 4.003 g.

#

1. The relative atomic mass of argon is 39.95 amu. What is the mass, in grams, of 3 moles of argon?

Confidence: 0 | 1 | 2 | 3 | 4

#

2. The relative atomic mass of oxygen is 16.00 amu, and that of hydrogen is 1.008 amu. How many moles of water are present in 1.000 kg of pure water?

Confidence: 0 | 1 | 2 | 3 | 4

#

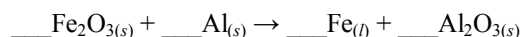
3. What is the mass of 2 moles of molecular oxygen?

Confidence: 0 | 1 | 2 | 3 | 4

3. The mole unit scales up the number of atoms by about 6 hundred thousand million million. It is directly proportional to the number of atoms (or molecules, or ionic units).

A balanced chemical reaction tells us how many of each type of atom (or molecule or ionic unit) react, to produce how many resulting atoms (or molecules or ionic units). That is, it describes the ratio of reactants and products. This ratio is the same when we measure the quantities of chemicals in moles.

- # 1. The thermite reaction is a reaction between iron oxide and aluminium. Balance its equation below.



Confidence: 0 | 1 | 2 | 3 | 4

- # 2. If we mix 200 g of Fe_2O_3 and 50 g of Al, we will have too much (an “excess”) of one of the reactants. Calculate the number of moles of each of these chemicals that we have.
(Atomic masses: O, 16.00 amu; Al, 26.98 amu; Fe, 55.85 amu)

Fe_2O_3 : _____ Al: _____ Confidence: 0 | 1 | 2 | 3 | 4

- # 3. Which of the reactants is in excess, and by how much?

Confidence: 0 | 1 | 2 | 3 | 4

4. Chemical concentration is measured in moles per litre (mol/L). We write [X] for “concentration of X”, and to calculate it, we divide the quantity of chemical (measured in moles) by the volume it is dispersed in (measured in litres).

- # 1. A particular sodium hydroxide solution has a concentration of 0.10 mol/L. How many moles of sodium hydroxide are present in 250 mL of this solution?

Confidence: 0 | 1 | 2 | 3 | 4

- # 2. What is the concentration (in mol/L) of a solution containing 10 g of table salt (NaCl) in 100 mL of solution? The atomic mass of Na is 22.99 amu; the atomic mass of Cl is 35.45 amu.

Confidence: 0 | 1 | 2 | 3 | 4

C2. Gases

The ideal gas equation relates pressure (P), volume (V), number of moles (n) and temperature (T) for any gas that behaves in an “ideal” fashion. It contains a factor R , the “ideal gas constant”, which makes the numbers the right size for the units that we use (pressure, kPa; volume, L; temperature, K). In this case, $R = 8.314 \text{ J/K/mol}$.

$$PV = nRT$$

- # 1. What is the volume of 1 mole of gas, at 101.325 kPa and 298 K?

Confidence: 0 | 1 | 2 | 3 | 4

C3. Atomic Structure and Periodicity

Electrons in atoms are arranged in shells and subshells. Larger shells contain more subshells than smaller shells. Specifically, the next larger shell will contain one additional type of subshell. Additional types of subshell come in odd number multiples.

We most commonly use numbers to refer to shells ($N = 1, 2, 3, \dots$) and letters to refer to subshells (s, p, d, f, \dots). The 1st shell contains a single s subshell. The 2nd shell contains a single s and three p subshells. The 3rd shell contains a single s , three p , and five d subshells.

- # 1. Predict how many f subshells the 4th shell would contain.

Confidence: 0 | 1 | 2 | 3 | 4

- # 2. Each subshell can contain up to 2 electrons. We can write electron configurations using superscripts to describe how many electrons are present in each type of subshell. For example, in zinc (element 30), with all its subshells filled so far, we write $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$.

Write the electron configuration for element 16, sulfur.

Confidence: 0 | 1 | 2 | 3 | 4

- # 3. The “ionisation energy” measures the amount of energy required to remove an electron from an atom. The first 8 ionisation energies for sulfur, in MJ/mol, are (in order): 1.006, 2.257, 3.367, 4.570, 7.019, 8.502, 27.113, 31.676. Explain this observed pattern.

Confidence: 0 | 1 | 2 | 3 | 4

C4. Structure and Bonding

Electron dot diagrams are a simple way to represent where we expect electrons from the outer shell of an atom to be. Usually, they are present in pairs, either as bonding pairs between two atoms, or as “lone pairs” that are not participating in bonding. For example, neon does not form bonds and has 4 lone pairs, while water has two bonding pairs and two lone pairs on the oxygen atom.



As a general rule (with many exceptions), try to distribute electrons so that each atom has 8 electrons in the outer shell (bonding pairs are counted for both atoms), unless the atom is hydrogen (which should have 2 electrons only). This rule applies to the cases below.

- # 1. Draw an electron dot diagram for carbon dioxide (both oxygen atoms will be bonded to the carbon atom, which is in the middle). Carbon starts with 4 electrons in its outer shell; oxygen with 6.



Confidence: 0 | 1 | 2 | 3 | 4

- # 2. Draw an electron dot diagram for carbon monoxide (one carbon atom bonded to one oxygen atom).



Confidence: 0 | 1 | 2 | 3 | 4

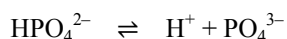
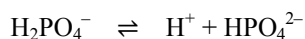
C5. Acids and Bases

1. One commonly used definition of acids and bases relies on the behaviour of H^+ ions, which are simply protons. A chemical that releases H^+ is an acid; a chemical that receives H^+ is a base.

Once an acid (we can represent this arbitrary acid as HA) releases a proton, it becomes a base (which we can write as A^-), because this new chemical can receive a proton to turn back into HA. These two chemicals (HA and A^-) together are called a “conjugate acid-base pair”.

If a chemical can act as both an acid and as a base, then it is called “amphoteric”. If it can do this under the definition involving protons, it is “amphiprotic”.

Phosphoric acid (H_3PO_4) is a triprotic acid: it has 3 hydrogen atoms that it can release as H^+ .



- # 1. Identify all conjugate acid-base pairs that appear in the equations above.

acid	base

Confidence: 0 | 1 | 2 | 3 | 4

- # 2. List all amphiprotic chemicals that appear in the equations above.

Confidence: 0 | 1 | 2 | 3 | 4

2. pH is the “hydrogen power”, which we calculate from the concentration of hydrogen ions in a solution.

$$pH = -\log_{10}[H^+]$$

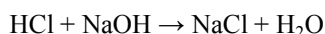
That is, pH is in fact a mathematical power (or logarithm), and 10 to the power of negative pH will give us the hydrogen ion concentration in a solution (in mol/L).

$$[H^+] = 10^{(-pH)}$$

- # 1. Calculate the pH of a hydrochloric acid (HCl) solution which has a concentration of 0.010 mol/L.

Confidence: 0 | 1 | 2 | 3 | 4

- # 3. Hydrochloric acid neutralises sodium hydroxide according to the following reaction.



- # 1. Calculate the pH of the final solution, after we mix 60 mL of 0.050 mol/L HCl solution, with 40 mL of 0.050 mol/L NaOH solution.

Confidence: 0 | 1 | 2 | 3 | 4

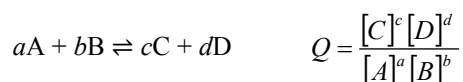
C6. Equilibrium

Chemical reactions are often incomplete, with the products tending to turn back into the reactants. As a reaction proceeds forwards, the concentration of products may rise, and the backward reaction will then speed up until it matches the rate of the forward reaction. This leads to chemical equilibrium.

- The ratio between reactants and products gives us an idea of where a chemical reaction is up to. We define this ratio as the “reaction quotient”, Q , in the following way.

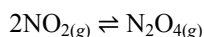
$$Q = \frac{\text{(concentrations of products, multiplied together)}}{\text{(concentrations of reactants, multiplied together)}}$$

The following example may be clearer, using arbitrary chemicals “A”, “B”, “C” and “D”, with coefficients a , b , c and d if they are required to balance the equation.



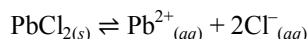
Do not include the concentrations of solids or liquids in the reaction quotient (use 1 instead if necessary). Only include the concentrations of dissolved substances, and gases.

- # 1. Write the reaction quotient expression for the following reaction.



Confidence: 0 | 1 | 2 | 3 | 4

- # 2. Write the reaction quotient expression for the following reaction.



Confidence: 0 | 1 | 2 | 3 | 4

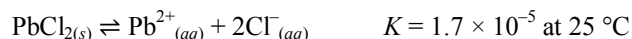
- As a chemical reaction reaches equilibrium, its reaction quotient will converge on a particular number. This number is different for different reactions, and for different temperatures.

We call this number the “equilibrium constant”, K .

$$K = (\text{value of } Q \text{ when the reaction is at equilibrium})$$

If we know what K is for a particular reaction at a particular temperature, then when we perform that chemical reaction, we can measure concentrations, and calculate Q to compare against K , to tell which way the reaction will proceed.

- # 1. Lead (II) chloride dissolves according to the following reaction.



In a particular solution of lead (II) chloride, the concentration of Pb^{2+} ions is 0.020 mol/L, and the concentration of Cl^{-} ions is 0.040 mol/L. Which is more likely to occur at this temperature: will more lead (II) chloride dissolve, or will lead (II) chloride tend to precipitate out of solution?

Confidence: 0 | 1 | 2 | 3 | 4

- # 2. Assume that Na^{+} and NO_3^{-} are very soluble and do not affect the other equilibria. We start with a solution of NaCl that has $[\text{Cl}^{-}] = 0.041$ mol/L. If we add $\text{Pb}(\text{NO}_3)_2$ to introduce $[\text{Pb}^{2+}]$, what concentration of Pb^{2+} must we achieve before PbCl_2 begins to precipitate?

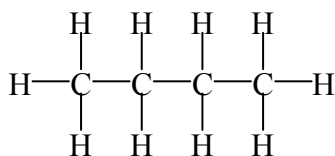
Confidence: 0 | 1 | 2 | 3 | 4

C7. Organic Chemistry

Organic chemistry is the chemistry of (most) carbon compounds. Here, we will deal with a few of these, and adhere to some specific rules: for the purposes of this test, carbon must form 4 bonds, oxygen must form 2 bonds, and hydrogen must form 1 bond.

#

- Isomers are chemicals with exactly the same atoms, but joined together in different arrangements. The structure of butane is drawn below, on the left. Draw an isomer of butane, to its right.

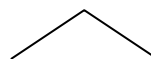
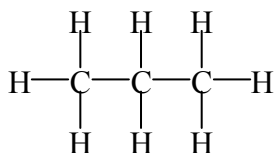


Confidence: 0 | 1 | 2 | 3 | 4

- Hydrocarbons contain only hydrogen and carbon. If all the carbon atoms in a hydrocarbon form a single straight (unbranched) chain, it is called a "straight-chain hydrocarbon", and named according to the number of carbon atoms in the chain. The name prefixes for 1 to 10 carbon atoms in a chain are given below.

atoms in chain	1	2	3	4	5	6	7	8	9	10
prefix	meth-	eth-	prop-	but-	pent-	hex-	hept-	oct-	non-	dec-

The example below is propane, drawn in full on the left, and as a "skeleton" on the right.



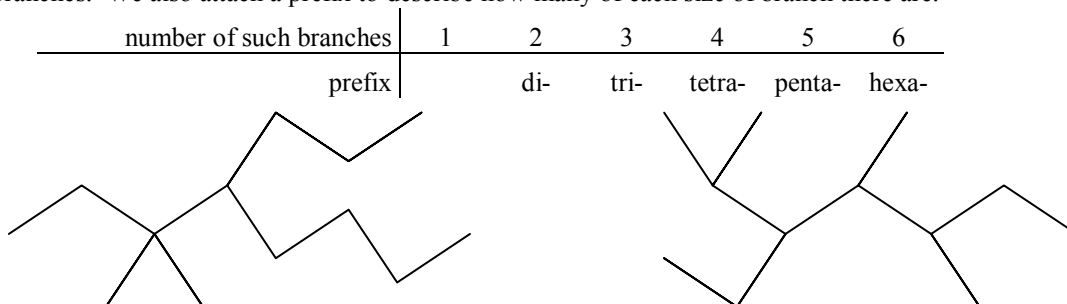
The skeleton form saves a lot of effort. To understand this notation, treat each corner or end as a carbon atom unless otherwise specified, and add hydrogen as required to make the correct total number of bonds.

#

- Draw hexane.

Confidence: 0 | 1 | 2 | 3 | 4

- For branched hydrocarbons, we name the longest straight chain, and then name the extra branches hanging off it. We should number positions along the main chain, to give the smallest total numbering for the branches. We also attach a prefix to describe how many of each size of branch there are.



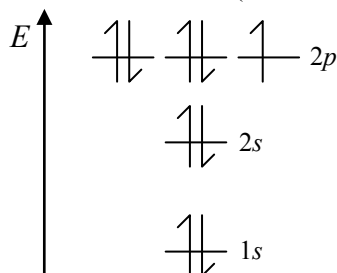
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- The example above on the left is 3,3-dimethyl-4-propyloctane. Name the hydrocarbon on its right.

Confidence: 0 | 1 | 2 | 3 | 4

C8. Covalent Bonding Models and d-orbital Chemistry

A more technical name for atomic subshells is “orbitals”. We can arrange atomic orbitals in order of energy on an “energy level diagram” like the one below (for fluorine, element 9).

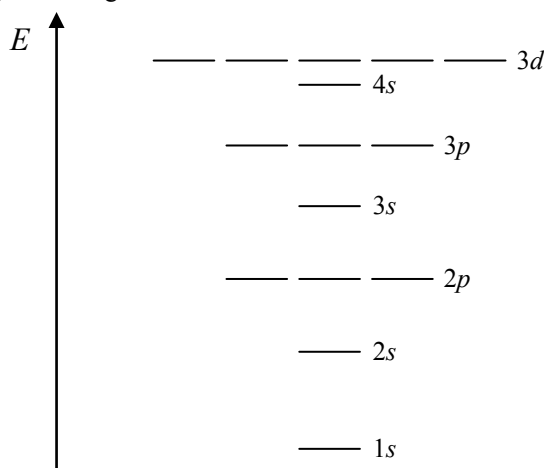


An atom in its ground state has all its electrons in the lowest possible energy levels. However, each orbital can only contain two electrons, with different “spin” (“up” or “down”). These are drawn as arrows, like above.

It takes some energy to pair electrons in the same orbital, so if multiple orbitals are close together in energy, it may be easier to place one electron in each of those orbitals first, before pairing them up.

#

1. Populate the energy level diagram with electrons for an atom of chromium (element 24) in its ground state.



Confidence: 0 | 1 | 2 | 3 | 4

C9. Electrochemistry

Chemistry can produce useful electrical power, a phenomenon we use in batteries. Electrochemistry depends on the movement of electrons from atom to atom.

1. We can assign atoms an “oxidation number” that describes whether an atom has acquired or lost electrons compared to its “neutral” state. To do this we follow certain rules, outlined below.
 - E1. Atoms in a pure element have oxidation number 0. (Each atom in O_2 or O_3 has oxidation number 0.) Also, a bond between atoms of the same type does not contribute to oxidation number.
 - E2. A monatomic ion has oxidation number equal to its charge. (Cl^- has oxidation number -1 .)
 - E3. The sum of oxidation numbers for a polyatomic species is the net charge on that species (0 for a neutral compound; the ion charge for a polyatomic ion).
 - E4a. Hydrogen has oxidation number $+1$ in all compounds except where it is the hydride ion (H^-), when it has oxidation number -1 .
 - E4b. Oxygen has oxidation number -2 in all compounds except where it is part of peroxide (O_2^{2-}), when it has oxidation number -1 .

Be aware that even in a single compound, different atoms of the same element may have different oxidation numbers, depending on what other atoms each one is attached to.

- # 1. Assign oxidation numbers to each atom in NH_4NO_3 , an ionic compound comprised of ammonium (NH_4^+) and nitrate (NO_3^-).

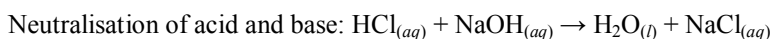
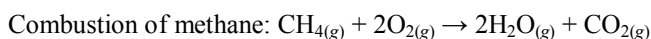
Confidence: 0 | 1 | 2 | 3 | 4

2. When electrons move from atom to atom, the atoms undergo a pair of processes: the atom that becomes more positive (oxidation number increased) is “oxidised”, while the atom that becomes more negative (oxidation number decreased, or reduced) is “reduced”. Remember that electrons have a negative charge, so more electrons means more negative.

Oxidation numbers describe which atoms the electrons are associated with, so we can tell if electrons are moving simply by watching for changes in oxidation numbers. If the oxidation numbers are changing, then an electrochemical reaction is occurring.

An oxidising agent, or “oxidant”, starts with a high oxidation number and causes another chemical to be oxidised (the oxidant itself loses the high oxidation number, and is reduced in the process). Conversely, a reducing agent, or “reductant”, starts with a low oxidation number and causes another chemical to be reduced (and itself increases in oxidation number, being oxidised in the process).

- # 1. For each reaction below, identify whether it is an electrochemical reaction.



Confidence: 0 | 1 | 2 | 3 | 4

3. For an electrochemical reaction, we can conceptually (and usually in practice) separate the oxidation reaction from the reduction reaction, as two “half-reactions”. This is useful, as it allows us to balance the half-reactions separately, as well as calculate the electrical potential we would expect from each reaction.

To balance electrochemical half-reactions, we follow certain rules, outlined below.

E1. Balance the interesting (non-O, non-H) atoms.

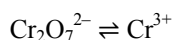
E2. To balance O, add $\text{H}_2\text{O}_{(l)}$.

E3. To balance H, add $\text{H}^+_{(aq)}$.

E4. To balance charge, add electrons (e^-).

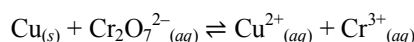
E5. (adjusts balancing in acid conditions to alkaline conditions) To use OH^- instead of H^+ , proceed as above, then add enough $\text{OH}^-_{(aq)}$ to both sides to neutralise all the $\text{H}^+_{(aq)}$ that appears, and then cancel out any $\text{H}_2\text{O}_{(l)}$ as appropriate.

- # 1. Balance the following half-reaction, which is the reduction of dichromate to chromium (III).



Confidence: 0 | 1 | 2 | 3 | 4

- # 2. Balance the following reaction by splitting it into half-reactions and combining them appropriately.



Confidence: 0 | 1 | 2 | 3 | 4

END