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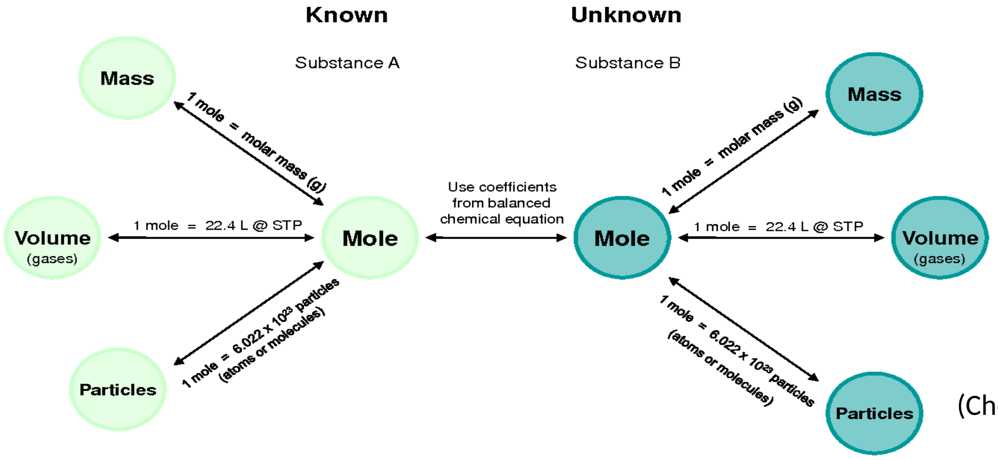
Chemical Reactions:

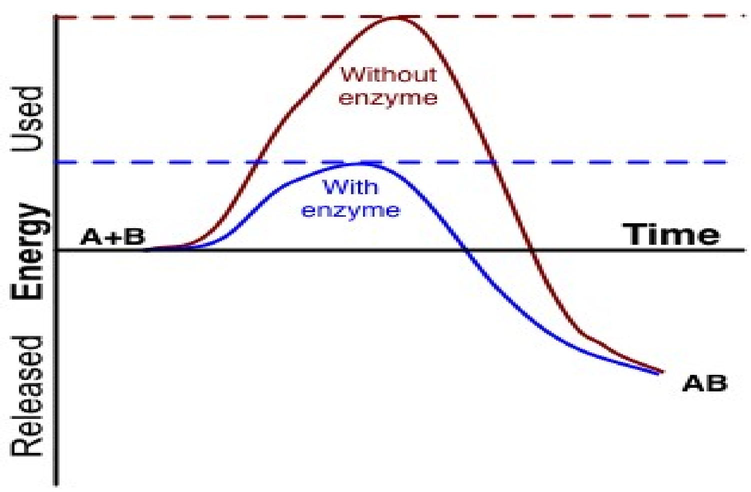
reactants, products and energy change

Stoichiometry Mole Island Diagram

When in doubt…convert to moles!



(Chemtutor n.d.)



(Biochemistry/Catalysis n.d.)

Week	Outcomes	References	Tasks
Term 2 Week 1-4	• chemical reactions can be represented by chemical equations; balanced chemical equations indicate the relative numbers of particles (atoms, molecules or ions) that are involved in the reaction • chemical reactions and phase changes involve enthalpy changes, commonly observable as changes in the temperature of the surroundings and/or the emission of light • endothermic and exothermic reactions can be explained in terms of the Law of Conservation of Energy and the breaking of existing bonds and forming of new bonds; heat energy released or absorbed by the system to or from the surroundings, can be represented in thermochemical equations • fossil fuels (including coal, oil, petroleum and natural gas) and biofuels (including biogas, biodiesel and bioethanol) can be compared in terms of their energy output, suitability for purpose, and the nature of products of combustion • the mole is a precisely defined quantity of matter equal to Avogadro's number of particles • the mole concept relates mass, moles and molar mass and, with the Law of Conservation of Mass; can be used to calculate the masses of reactants and products in a chemical reaction • empirical formula can be determined using percentage composition, mass composition and combustion data • the limiting reagent in a chemical reaction can be determined using masses and moles of reactants	Lucarelli p 45 Set 8 Lucarelli p 100 -108 Set 22 Q1-9 Lucarelli p 77 - 80 Set 16 Q1-5 Lucarelli p 81- 82 Set 17 Q1-12 Lucarelli p 86 - 88 Set 19 Q1-6	STAWA Experiment 21 STAWA Experiment 35

Week	Outcomes	References	Tasks
5-7	REVISION AND EXAMS		
			Task 8: Extended response 2- Energy and CO_2 output for fossil fuels and biofuels **Task 9: Test- Chemical Reactions: Reactants, Products and Energy Change**

Write balanced chemical equations of the following for more practice.

1. Hydrogen + oxygen → water
2. Carbon + oxygen → carbon dioxide
3. Sodium + chlorine → sodium chloride
4. Potassium + oxygen → potassium oxide
5. Magnesium + oxygen → magnesium oxide
6. Magnesium + hydrochloric acid → magnesium chloride + hydrogen
7. Ammonia + water → ammonium hydroxide
8. Zinc + hydrochloric acid → zinc chloride + hydrogen
9. Sodium + sulfuric acid → sodium sulfate + hydrogen
10. Calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide
11. Hydrogen + nitrogen → ammonia
12. Aluminium + oxygen → aluminium oxide
13. Calcium oxide + water → calcium hydroxide
14. Calcium hydroxide + carbon dioxide → calcium carbonate + water
15. Ammonia + sulfuric acid → ammonium sulfate
16. Silver nitrate + magnesium chloride → silver chloride + magnesium nitrate
17. Potassium nitrate → potassium nitrite + oxygen
18. Sodium hydroxide + carbon dioxide → sodium carbonate + water
19. Potassium hydrogen carbonate → potassium carbonate + water + carbon dioxide
20. Zinc + silver nitrate → zinc nitrate + silver

ENDOTHERMIC AND EXOTHERMIC REACTIONS (pg 101)

All chemical reactions involve energy changes. Energy can be either gained or released by the reactants or products. This follows the law of conservation of energy which states that energy cannot be created or destroyed, merely converted from one form to another.

Terms:

- **Enthalpy (H):** the total energy (both chemical stored potential and kinetic energy) present in a substance.
- **ΔH (enthalpy change):** difference between enthalpy of products and the enthalpy of reactants. The difference between the energy gained (absorbed/ taken in/ required) and the energy released. $\Delta H = H_{products} - H_{reactants}$
- **System:** the collection of atoms or molecules involved in a chemical reaction e.g. the ions in a precipitation reaction.
- **Surroundings:** anything around the system, but is not part of the system e.g. the solvent in a precipitation reaction, the test tube, your hand.

<u>ENERGY PROFILE DIAGRAMS FOR EXOTHERMIC AND ENDOTHERMIC REACTIONS</u>

The release or gain of energy can be represented in diagrams where reaction time is plotted against enthalpy. Note that total energy is always conserved, despite the change in energy of a substance as shown in these diagrams – copy from Essential Chemistry pg 114

Exothermic Reactions

1. More energy is released in forming bonds than is required to break the initial bonds, so the reaction releases energy to the surroundings. (Bond forming releases energy, bond breaking requires energy. The stronger the bonds that are formed, the greater the energy released in forming them.)
2. Enthalpy of products is less than reactants.
3. Reactants have a higher chemical energy than the products.
4. ΔH is negative.
5. Enthalpy decreases for the system and increases for the surroundings.
6. Energy is released from the system to the surroundings in the form of heat, light, etc.
7. The surroundings will gain heat.

<u>Examples of exothermic processes</u>

- Combustion: $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)} + 803\ kJ$ **OR**

  $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}\quad \Delta H = -803\ kJ$
- Reactions involving single atoms bond together: $2I_{(g)} \rightarrow I_{2(g)} + 214\ kJ$
- Reactions in which a positive ion gains an electron: $Na^+_{(g)} + e^- \rightarrow Na_{(g)}$
- Condensation and solidification (freezing) phase changes: $I_{2(g)} \rightarrow I_{2(s)} + 62\ kJ$

  $H_2O_{(g)} \rightarrow H_2O_{(l)}\quad \Delta H = -44\ kJ$
- Respiration: $C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$
- Heat packs

  – non-reversible: $4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}\quad \Delta H = -1654\ kJ$ **OR**

  – reversible: $NaCH_3COO_{(aq)} + 3H_2O_{(l)} \leftrightharpoons NaCH_3COO.3H_2O_{(s)}$
- Crystallisation

Phase changes are physical processes that involve the breaking and forming of bonds but of weak bonds only like weak intermolecular forces between iodine molecules during sublimation. By comparison chemical changes involve the breaking and forming of strong intramolecular bonds like covalent bonds in the iodine molecule. For this reason, phase changes usually involve relatively small amounts of energy compared to chemical changes.

In general, chemical processes involve greater enthalpy changes than physical processes. This is due to the fact that the intermolecular forces are usually weaker than the intramolecular forces.

Endothermic reactions

1. More energy is required in breaking the initial bonds than is released in forming bonds so the reaction absorbs energy from the surroundings. (Bond forming releases energy, bond breaking requires energy. The stronger the bonds that are broken, the greater the energy required to break them)
2. Enthalpy of products is greater than reactants.
3. Reactants have a lower chemical energy than the products.
4. ΔH is positive.
5. Enthalpy increases for the system and decreases for the surroundings.
6. Energy is absorbed from the surroundings to the system.
7. The surroundings will lose heat.

<u>Examples of endothermic processes:</u>

- Reactions involving a molecule breaking up: $F_{2(g)} \rightarrow 2F_{(g)}$ $\Delta H = +158$ kJ
- Reactions in which an atom or ion loses an electron: $K_{(g)} \rightarrow K^{+}_{(g)} + e^{-}$
- Vaporisation and liquefaction phase changes.
- Photosynthesis: $6CO_{2(g)} + 6H_2O_{(l)} \rightarrow C_6H_{12}O_{6(aq)} + 6O_{2(g)}$
- Cold packs (older style involving ammonium nitrate dissolving in water)
  - $NH_4NO_{3(s)} + 126\text{ kJ} \rightarrow NH_4^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)}$ **OR**
  - $NH_4NO_{3(s)} \rightarrow NH_4^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)}$ $\Delta H = +126$ kJ

Comparing fossil fuels: emissions and fuel values.

Fuel values (sometimes called heating values): compare the energy from the complete combustion of equal masses or volumes of different fuels. The greater the value the greater the energy available from a given mass. Units in kJ g^{-1}, MJ kg^{-1}, MJ L^{-1}

$x = \frac{\Delta Hc}{M}$ Where ΔHc = standard heat of combustion of 1 mol of the fuel. M = molar mass

Carbon emissions: Combustion of fuels release carbon dioxide, which is a known greenhouse gas, into the atmosphere.

Carbon emission values: compare the mass of carbon dioxide a given fuel produces with the amount of energy released. Units in g MJ^{-1}.

Fill in the values for the table below from Essential Chemistry pg 105.

	Heating value $MJ.kg^{-1}$	**Carbon emissions $g(CO_2)MJ^{-1}$**
Coal		
Natural gas		
LNG		
lPG		
Petrol		
Diesel		

Biofuels

Biofuels are fuels that are produced from biodegradeable materials such as crops rather than from fossil fuels. Examples of biofuels include: bioethanol, biogas and biodiesel.

Advantages of biofuels are that they: are made from a renewable resource, they have lower carbon emissions and they have extremely low sulphur content meaning there is no SO_2 formation that can lead to acid rain.

Bioethanol is produced from the fermentation of plant sugars (such as in wheat and sugar cane) to ethanol by yeast.

Biodiesel is produced by the transesterification of oilseed crops.

.*Significant Figures and Rounding*

The rules for determining the number of significant figures are (Essential Chemistry pg 102):

Empirical formula

Empirical formula can be calculated using:

a) Mass data
b) Percentage composition
c) Combustion data

a) A compound was analysed and found to contain by mass 4.659 g silver and 0.347 g oxygen. Determine its empirical formula.

b) An unidentified compound was analysed in the laboratory and found to have the following percentage composition by mass: Carbon 26.09%, Hydrogen 4.35%, Oxygen 69.56%. What is the empirical formula?

Also remember how to calculate percentage composition. Calculate the percentage composition of each element in $Al(NO_3)_3$.

c) A white crystalline solid of molecular formula $C_xH_yO_z$ was isolated from certain fruits. When 0.682 g of the compound was combusted with oxygen gas, it produced 0.968 g of carbon dioxide gas and 0.594 g of water. Determine the empirical formula and, hence, the molecular formula of the substance if it has a molecular mass of 62.3 g mol^{-1}.

Avogadro's number (N)

Avogadro's number is **6.022×10^{23}** number of particles.

For a pure substance, this is the number of particles (i.e. atoms, molecules or formula units) in **one mole** of that substance.

The mole (mol)

A mole of any substance is equivalent to 6.022×10^{23} particles of that substance. For a pure substance, this will be equal to its molar mass in grams.
Examples:
a) 1 mol of Mg contains 6.022×10^{23} atoms of Mg and has a mass of 24.31 g.
b) 1 mol of carbon dioxide gas (CO_2) contains 6.022×10^{23} molecules of CO_2 and has a mass of 44.01 g.
c) 1 mol of NaCl contains 6.022×10^{23} formula units of NaCl and has a mass of 58.44 g.
The relationship between the number of moles of a substance (n), its mass (m) and the molar mass (M) is:

$$\mathbf{n = m/M}$$

Stoichiometry

A balanced chemical equation is based on the principle of the **Law of Conservation of Mass.**

The law states that in a chemical reaction, matter cannot be created or destroyed. The mass of the reactants before the reaction will equal the mass of the products after the reaction, or the number of atoms of each element is the same on both sides of the equation.

The balanced chemical equation shows the number of particles of each species that reacts or is produced. Because a mole is directly proportional to the number of particles then the coefficients in a balanced equation give the ratio in which each species reacts or is produced. We call this the mole ratio and it can be calculated using the following formula:

$$\mathbf{n_u = n_k \times u/k}$$

where: n_U = number of moles of unknown species
n_k = number of moles of known species
u= coefficient of unknown species
k = coefficient of known species

Limiting reagent

We have learned how to work out the moles and masses which react in a chemical equation when one of the reactants is already known. However, many times when two reactants are added together, there is not enough of one to completely react with the other reactant.

Eg If you had 32 g of a magnesium metal and only had a very small volume of dilute acid for it to react with then not all of the magnesium will react and you would have some left over at the end.

In this case, the limiting reagent is the acid and the excess reagent is the magnesium.

The limiting reagent will affect the amount of the other reactants used and the amount of products produced.

Eg. Which reactant would be the limiting reagent in the following? Explain why.

$$2Mg \quad + \quad O_2 \quad \rightarrow \quad 2MgO$$

a) 1 mol of each

b) 2 mol Mg, 1 mol O_2

c) 1 mol Mg, 2 mol O_2

d) 3.7 mol Mg, 2.4 mol O_2

Mole Constituents

1 mol HNO_3 contains: _________________________ mol H atoms _________________________ mol N atoms _________________________ mol O atoms	1 mol $Cu(NO_3)_2$ contains: _________________________ mol Cu atoms _________________________ mol N atoms _________________________ mol O atoms
1 mol sulphur dioxide contains: _________________________ mol S atoms _________________________ mol O atoms	1 mol barium nitrate contains: _________________________ mol Ba atoms _________________________ mol N atoms _________________________ mol O atoms
1 mol $(NH_4)_2SO_4$ contains: _________________________ mol N atoms _________________________ mol H atoms _________________________ mol S atoms _________________________ mol O atoms	5.6 mol $(NH_4)_2SO_4$ contains: _________________________ mol N atoms _________________________ mol H atoms _________________________ mol S atoms _________________________ mol O atoms
3 mol KNO_3 contains: _________________________ mol K atoms _________________________ mol N atoms _________________________ mol O atoms	7.2 mol $Ca(OH)_2$ contains: _________________________ mol Ca atoms _________________________ mol O atoms _________________________ mol H atoms
0.24 mol $ZnCl_2$ contains: _________________________ mol Zn atoms _________________________ mol Cl atoms	0.8 mol $PbSO_4$ contains: _________________________ mol Pb atoms _________________________ mol S atoms _________________________ mol O atoms

1. Complete the table below:

Name	Symbol/ Formula	Molar mass (M) (g mol^{-1})	Number of moles (n) (mol)	Mass (m) (g)	Number of particles (N)	Mass of one particle (g)
	Zn				5.15×10^{23} atoms	
Sulfur trioxide			2.74			
Lead (II) nitrate				186.35		

2. Calculate (showing all working) the **number of moles** of the following:

    a) 876.2g of $MnBr_2$
    b) 93.4g of $Cr(NO_2)_3$

3. Calculate (showing all working) the **mass** of:

    a) 5.000 moles of aluminium bromide($AlBr_3$).
    b) 10.00 moles of calcium nitrate ($Ca(NO_3)_2$).

4. Nitrogen gas reacts with hydrogen gas to produce ammonia gas according to the following reaction:

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

    If 2 moles of nitrogen gas reacts, calculate (showing all working):

    a) The number of moles of hydrogen gas reacted.
    b) The number of moles of ammonia gas produced.

5. 3.5 mole of tin reacts according to the following reaction:

$$3\,Sn(s) + 16HNO_3(aq) \rightarrow 3\,Sn(NO_3)_4(aq) + 4NO(g) + 8\,H_2O(l)$$

    Calculate:
    a) The mass of nitric acid that reacts.
    b) The mass of water produced.

6. 32.7 g of aluminium is reacted with 131 g of chlorine to produce aluminium chloride.

Determine:

a) The limiting reagent.
b) The mass of aluminium chloride produced.
c) The mass of excess reagent left over.

7. 0.035 moles of HCl reacts with 2.5 g of $CaCO_3$. How many moles of CO_2 are formed?

8. In **3 mol of $Ca_3(PO_4)_2$**, what are the number of moles of the following;

Particle	Ca^{2+} ions	PO_4^{3-} ions	P atoms	O atoms
n				

9. Calculate the number of moles of O atoms in 986.5g of $Fe(HSO_4)_3$

Bibliography

Chemtutor n.d.: , (Chemtutor n.d.),

Biochemistry/Catalysis n.d.: , (Biochemistry/Catalysis n.d.),