WHAT’S THE MATTER?

Once again, we start with some Year 11 work.

Here is the 3AB outline of this topic.

Macroscopic properties of matter

- Interpret observations, such as the colour changes of physical and chemical systems at equilibrium
- Use observable properties, such as the colour of ions, to help predict and explain the formation of products in chemical processes (see data sheet)
- Use the Kinetic Theory to explain the concept of absolute zero

Solutions

- Apply the solubility rules to predict if a precipitate will form when two dilute ionic solutions are mixed (see data sheet)
- Perform concentration calculations (mol L^{-1}, g L^{-1} ppm, percentage composition)
- Calculate the concentration of ions in solution for strong electrolytes
- Perform the calculation of concentration and volume involved in the dilution of solutions and the addition of solutions

Chemical Reactions

Reactions, equations and stoichiometry

- Describe, write equations for and interpret observations for the following reaction types:
  - ■ Precipitation
  - ■ Solvation of ions in aqueous solution
  - ■ Physical and chemical equilibrium
- Write ionic equations appropriate to the chosen context using ions in appendix 1.
- Perform calculations involving
  - ■ Conversion between Celsius and Kelvin temperature scales
  - ■ Mass, molar mass, number of moles of solute, concentration and volume of solution and gas volume using PV=nRT
  - ■ Percentage purity of reactants or percentage yield in industrial processes
  - ■ A limiting reagent, including:
    - o Identification of limiting reagents
    - o Calculation of excess reagents

Applied Chemistry

- Write the chemical formulae for molecular compounds based on the number of atoms of each element present as inferred from the systematic names

Write the molecular formulae of commonly encountered molecules that have non-systematic names

This part of the course is largely revision, with some extensions of ideas you are already familiar with.

Here is a list of tasks in what I think is a logical sequence. The list includes some reading, Review Exercises, Problem Sets and experiments.

Work through the list spending most time on the tasks that challenge you the most.

The Kinetic Theory and Absolute Zero

STAWA Exp't 1 page 28

Do not attempt this experiment. Instead, use this data to help you answer the questions.

This experiment demonstrates why we always use the Kelvin scale when associating temperature to the other physical dimensions of gases (P and V). The explanation of the results is based on collision theory.

Here is some data for this experiment. Complete the table, consider the information it provides and then answer the Process Questions from page 30.

Temperature ^{o}C	Temperature K	Pressure kPa
-10		94
0		98
25		112
50		115
85		128

- **Apply the solubility rules to predict if a precipitate will form when two dilute ionic solutions are mixed (see data sheet)**

QuickTime™ and a
TIFF (LZW) decompressor
are needed to see this picture.

Chemistry for WA Chapter 5.1 and exercise 5.1

Chapter 5.2 and exercise 5.2

STAWA Investigation 1 page 41 - Identify the unknowns –

Different substances dissolve in water to different extents.

Apply the solubility rules to identify the unknown solutions in

Sets 1 and 2

- **Perform concentration calculations (mol L^{-1}, g L^{-1} ppm, percentage composition)**
- **Calculate the concentration of ions in solution for strong electrolytes**
- **Perform the calculation of concentration and volume involved in the dilution of solutions and the addition of solutions.**
- **Calculate the concentration of ions in solution for strong electrolytes**
- **Perform the calculation of concentration and volume involved in the dilution of solutions and the addition of solutions**

QuickTime™ and a
TIFF (LZW) decompressor
are needed to see this picture.

Chemistry for WA Chapter 4.2 and exercise 4.2

Chapter 4.4 and exercise 4.4

4.4 Solution concentrations

The concentration of a solution refers to the quantity of solute dissolved in a particular quantity of solution. The mass or volume of the solution includes the mass or volume of both the solute and the solvent.

Four methods commonly used to express solution concentration are:

- mole per litre, mol L^{-1}
- gram per litre, g L^{-1}
- parts per million, ppm
- percentage composition by mass.

Chapter 4.5 and exercise 4.5

4.5 Diluting and mixing solutions

When more solvent is added to a solution, the concentration of the solution decreases. In this dilution process, the volume of the solution is increased, but the amount of solute remains constant. That is:

mole of solute in concentrated solution = mole of solute in dilute solution

A method sometimes used to calculate the new concentration of a diluted solution is based on the following reason.

If mole of solute in the initial concentrated solution
= mole of solute in the final diluted solution

and because

mole of solute = concentration of solution × volume of solution

then it follows that:

concentration of initial solution × initial volume
= concentration of final solution × final volume

or

$$c_1V_1 = c_2V_2$$

where c_1 and V_1 refer to the concentrated solution before dilution and c_2 and V_2 refer to the diluted solution.

Questions page 118: 3, 4, 9, 10, 13, 14, 16, 18

Perform calculations involving

- **Conversion between Celsius and Kelvin temperature scales**
- **Mass, molar mass, number of moles of solute, concentration and volume of solution and gas volume using PV=nRT**
- **Percentage purity of reactants or percentage yield in industrial processes**
- **A limiting reagent, including:**
    - **Identification of limiting reagents**
    - **Calculation of excess reagents**

Chemistry for WA Chapter 5.3 and exercise 5.3

5.3 Stoichiometry

A balanced chemical equation shows the relationships between the number of particles of reactants and products in a chemical reaction. For example, the following balanced equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

shows that 1 molecule of propane (C_3H_8) will react with 5 molecules of oxygen to form 3 molecules of carbon dioxide and 4 molecules of water.

This stoichiometric ratio (the ratio shown in a balanced equation) between the reactants and the products they will form is also correct if the amount of each substance is expressed in mole. That is:

1 mol of C_3H_8 reacts with 5 mol of O_2 to form 3 mol of CO_2 and 4 mol of H_2O

or

2 mol of C_3H_8 reacts with 10 mol of O_2 to form 6 mol of CO_2 and 8 mol of H_2O

or

0.1 mol of C_3H_8 reacts with 0.5 mol of O_2 to form 0.3 mol of CO_2 and 0.4 mol of H_2O

Chapter 5.4 and exercise 5.4

5.4 Gases and stoichiometry

Gas laws

During the 17th and 18th centuries, various relationships between the pressure, volume, temperature and amount in mole of gases were investigated. The following laws resulted from these investigations:

- *Boyle's law*: If the temperature of a gas is kept constant, the volume of a given mass of gas is inversely proportional to its pressure.

  That is, $V \propto \frac{1}{P}$ where the volumes and pressures are measured at a constant temperature and for a constant mass of gas.
- *Charles' law*: At constant pressure, the volume of a given mass of gas is directly proportional to its temperature on the kelvin scale.
  This law can be represented as $V \propto T$ for a constant mass of gas at a constant pressure, and where the temperature is in kelvin.
  (kelvin temperature = Celsius temperature + 273.1)
- *Avogadro's hypothesis*: Equal volumes of all gases, measured at the same temperature and pressure, contain equal numbers of particles (or amount in mole of particles).
  That is, $V \propto n$ where the volumes are measured at the same temperature and pressure.

The two laws and Avogadro's hypothesis can be summarised in the ideal gas equation. This equation gives the relationship between the volume, pressure, temperature and amount in mole of a sample of a gas.

$$PV = nRT$$

where P is the pressure, V is the volume, n is the amount in mole, R is a constant (the universal gas constant) and T is the temperature in kelvin.

The universal gas constant, R, has a value of 8.315 J mol^{-1} K^{-1}, when P is measured in kPa, V in L and T in K.

At times it is necessary to convert pressures to the units required in a particular equation. Several useful conversions are:

1 atmosphere = 760 mmHg = 101.3 kPa

STAWA Problem Set 1 Gases p 47

Chapter 5.5 and exercise 5.5

5.5 Limiting reagents

Sometimes two reactants are mixed in 'non-stoichiometric' amounts, resulting in one of these reactants not being completely consumed in the reaction. The reactant that is completely consumed is referred to as the limiting reagent and the reactant that remains after the reaction is the excess reagent. The amount of limiting reagent present in the reaction mixture determines the amount of products formed during the reaction.

For example when nitrogen reacts with hydrogen in the Haber process, the following reaction occurs:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

This equation shows that hydrogen and nitrogen react with a 3:1 molar amount. The data given in Table 5.4 illustrates what happens when different mixtures of the reactants are reacted together (assuming the reaction is 100% complete).

TABLE 5.4 EXAMPLES OF LIMITING REAGENT MIXTURES FOR THE REACTION $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Initial amount of N_2	Initial amount of H_2	Limiting reagent	Excess reagent	Amount of NH_3 formed	Amount of excess reagent left after the reaction
1 mol	3 mol	neither	neither	$1 \times 2 = 2$ mol	0 mol
2 mol	3 mol	H_2	N_2	$\frac{2}{3} \times 3 = 2$ mol	$2 - 1 = 1$ mol of N_2
1 mol	4 mol	N_2	H_2	$1 \times 2 = 2$ mol	$4 - 3 = 1$ mol of H_2
2 mol	8 mol	N_2	H_2	$2 \times 2 = 4$ mol	$8 - 6 = 2$ mol of H_2
4 mol	6 mol	H_2	N_2	$\frac{2}{3} \times 6 = 4$ mol	$4 - 2 = 2$ mol of N_2
9 mol	9 mol	H_2	N_2	$\frac{2}{3} \times 9 = 6$ mol	$9 - 3 = 6$ mol of N_2

In stoichiometric calculations where the amounts of two reactants have been given, i.e. limiting reagent problems, the first step is to determine which reactant is the limiting reagent. It is then the amount of limiting reagent that is used to determine the amount of product produced.

STAWA Problem Set 2 Limiting Reagents p 50

STAWA Investigation 3 page 46 – Determine the limiting reagent –
Gravimetric analysis (analysis using masses) can be used to determine the limiting reagent in many reactions.
The limiting reagent determines how much product is made. It is the chemical that gets used up.

Chapter 5.6 and exercise 5.6

5.6 Stoichiometry and percentages

Often problems either give information concerning the percentage purity of a particular substance or the percentage yield of a reaction, or they ask for percentages such as these to be calculated.

The percentage purity, by mass, of a substance can be calculated using the formula:

$$\text{percentage purity of the material} = \frac{\text{mass of pure substance}}{\text{mass of impure substance}} \times 100$$

The percentage yield of a reaction is a measure of how much product was actually produced compared to the theoretical amount expected to be produced. This theoretical amount can be calculated using the stoichiometric ratio given in the balanced chemical equation. One equation that can be used to determine the percentage yield of a reaction is:

$$\text{percentage yield of a reaction} = \frac{\text{mass of product actually obtained}}{\text{theoretical mass of product expected}} \times 100$$

The amount in mole of the product formed, compared to the theoretical amount in mole, can also be used to determine the percentage yield, as can the actual and theoretical volumes (measured at the same temperature and pressure) of a gaseous product.

STAWA Problem Set 5 Percentage Composition and Yield p57

STAWA Investigation 2 page 45 – Determine the formula-

Determining the empirical formula of a compound is an important analysis technique. This investigation involves several steps. Make sure you know where you are going before you set off!

Questions page 148: 1, 2, 3, 4, 5, 6, 7, 11, 14, 16