JABchem

Name: …………………………………………………………………………………….	
Class: ………………	Teacher: ………………………………

Self-Evaluation

Higher Chemistry

Unit 3

Chemistry In Society

Section	Title	Completed
3.1	Getting The Most From Reactants	
3.2a	Controlling The Rate – Collision Theory	
3.2b	Controlling The Rate – Reaction Pathways	
3.2c	Controlling The Rate – Kinetic Energy Distribution	
3.3a	Chemical Energy - Enthalpy	
3.3b	Chemical Energy - Hess's Law	
3.3c	Chemical Energy - Bond Enthalpies	
3.4	Equilibria	
3.5a	Chromatography	
3.5b	Volumetric Titrations	

Higher Chemistry Self-Evaluation
Unit 3.1 Getting The Most From Reactants

		Page	Red	Amber	Green
120	Industrial processes are designed to maximise profits and minimise the impact on the environment		☹	😐	☺
121	Factors influencing industrial process design include: a) availability, sustainability and cost of freestocks b) opportunities for recycling c) energy requirements d) marketability of by-products e) product yield		☹	😐	☺
122	Environmental considerations include: a) minimising waste b) avoiding the use or production of toxic substances c) designing products which will biodegradable, if appropriate		☹	😐	☺
123	Chemical reactions are written with formulae can show the mole ratio of reactants and products: $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$ 1mol, 3mol, 2mol		☹	😐	☺
124 129a	1 mole of a substance is equal to the gram formula mass and is calculated from relative atomic masses. Calculations can be performed using gram formula mass e.g. Calculate the gfm of calcium nitrate: 1mol = gfm $Ca(NO_3)_2$ = (1x40.1) + (2x14) + (6x16) = 40.1 + 28 + 96 = 164.1g		☹	😐	☺
125 129b	Calculations turning masses into number of moles (and vice versa) require the gfm: Calculate the number of moles in 0.328g of calcium nitrate? gfm $Ca(NO_3)_2$ =164.1g mol^{-1} no. of mol = $\frac{\textbf{mass}}{\textbf{gfm}} = \frac{0.328}{164.1}$ = 0.002mol Calculate the mass of 0.05mol of calcium nitrate? gfm $Ca(NO_3)_2$ =164.1g mol^{-1} **mass** = **no.** of mol x **gfm** = 0.05x164.1 = 8.21g Calculation involving masses and number of moles can also involve the mole ratio in balanced equations: e.g. *calculate the mass of carbon dioxide produced if 5g of calcium carbonate reacts with excess HCl* gfm $CaCO_3$ = (1x40.1)+(1x12)+(3x16) = 40.1+12+48 = 100.1g gfm CO_2 = (1x12)+(2x16) = 12+32 = 44g no. of mol = $\frac{\textbf{mass}}{\textbf{gfm}} = \frac{5}{100.1}$ = 0.05mol $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ 1mol ($CaCO_3$), 1mol (CO_2) 0.05mol ($CaCO_3$), 0.05mol (CO_2) **mass** = **no.** of mol x **gfm** = 0.05 mol x 44 g mol^{-1} = 2.2g		☹	😐	☺
126 129d	Calculations can be performed using volumes and concentrations e.g. *Calculate the mass of calcium carbonate required to completely react with 80cm³ of 0.1mol l⁻¹ hydrochloric acid.* **no.** of mol HCl = **volume** x **concentration** = 0.08litres x 0.1mol l^{-1} = 0.008mol $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ 1mol ($CaCO_3$), 2mol (HCl) 0.004mol ($CaCO_3$), 0.008mol (HCl) **gfm** $CaCO_3$ = (1x40.1)+(1x12)+(3x16) = 40.1+12+48 = 100.1g **mass** = **no.** of mol x **gfm** = 0.004mol x 100.1g mol^{-1} = 0.4004g e.g. *Calculate the concentration of a solution when 5.85g of NaCl is dissolved in 50cm³ water.* Calculate the gfm of NaCl: Na 1 x 23 = 23; Cl 1 x 35.5 = 35.5; gfm = 58.5g Calculate number of moles of NaCl: $n = \frac{m}{gfm} = \frac{5.85}{58.5}$ = 0.1mol Calculate the concentration: $c = \frac{n}{V} = \frac{0.1mol}{0.05litres}$ = 2 mol l^{-1}		☹	😐	☺
127 129f	The molar volume is the volume occupied by one mole of gas at a certain temperature and pressure. • The molar volume (in l mol^{-1}) is the same for all gases at the same temperature and pressure • The volumes of reactant and product gases can be calculated from the number of moles of each reactant and product. e.g. *Calculate the final volume and composition of the mixture produced when 100cm³ of ethane is completely burned in 500cm³ of oxygen.* $C_2H_{6(g)} + 3\frac{1}{2}O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)}$ 1mol, 3.5mol, 2mol, 3mol 1vol, 3.5vol, 2vol, negligible volume 100cm³, 350cm³, 200cm³, -		☹	😐	☺

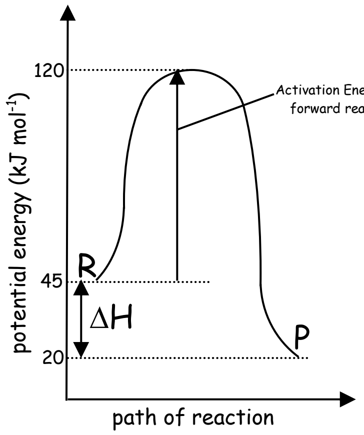
	(+150cm^3 O_2 leftover) Final Volume = 350cm^3 (200cm^3 CO_2 + 150cm^3 O_2)				
128 129e	The volume of a gas can be calculated from the number of moles and vice versa using the molar volume e.g. *Calculate the volume of 0.8g of oxygen gas if molar volume = 24* litres mol^{-1} $\mathbf{n}\text{o. of mol } O_2 = \frac{\textbf{mass}}{\textbf{gfm}} = \frac{0.8\text{g}}{32\text{g mol}^{-1}} = 0.025\text{mol}$ **V**olume = **n**o. of mol x **M**olar **V**olume = 0.025mol x 24litres mol^{-1} = <u>0.6litres</u>		☹	😐	☺
131	I can use a balanced equation to work out the reactant in excess and therefore the limiting reactant, for a chemical reaction. e.g. Calculate which reactant is in excess when 2g of hydrogen reacts with 8g of oxygen. **gfm** H_2 = 2g ∴ n=$^{mass}/_{gfm}$ = $^2/_2$ = 1mol **gfm** O_2 = 32g ∴ n=$^{mass}/_{gfm}$ = $^8/_{32}$ = 0.25mol $2H_2 + O_2 \longrightarrow 2H_2O$ 2mol 1mol 1mol 0.5mol (required) but only 0.25mol O_2 of oxygen available ∴ hydrogen is in excess ∴ oxygen is the limiting reactant as it runs out $2H_2 + O_2 \longrightarrow 2H_2O$ 2mol 1mol 0.5mol 0.25mol 0.5mol of H_2 needed to react all O_2 and 1mol of H_2 available ∴ hydrogen is in excess ∴ oxygen is the limiting reactant as it runs out		☹	😐	☺
132	In order to ensure that a costly reactant is converted in a product, an excess of the less expensive reactant(s) is used		☹	😐	☺
133 134 135	For a particular set of conditions, the percentage yield provides a measure of the degree to which the limiting reagent is converted into the desired product. • The *theoretical yield* is the quantity of desired product obtained, assuming full conversion of the limiting reagent, as calculated from the balanced equation. • The *actual yield* is the quantity of the desired product formed under the prevailing reaction conditions		☹	😐	☺
130a 136	Percentage yield is a measure of how much of a product is obtained compared to the amount expected if there was complete conversion. $\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$		☹	😐	☺
137	Percentage yields can be calculated from balanced equations and masses of reactants and products: e.g. Calculate the %yield of ester if 2g of methyl ethanoate is formed when 1.6g of methanol is used. methanol + ethanoic acid → methyl ethanoate + water $CH_3OH + CH_3COOH \rightarrow CH_3OCOCH_3 + H_2O$ 1mol 1mol 32g 74g 1.6g 74g x $^{1.6}/_{32}$ Theoretical = 3.7g $\%\text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100$ $= \frac{2}{3.7\text{g}} \times 100$ $= 54.1\%$		☹	😐	☺
138	Given costs for the reactants, a percentage yield can be used to calculate the cost of reactant(s) required to produce a given mass of product		☹	😐	☺
130b 139	Atom economy measures the proportion of the total mass of all starting materials successfully converted into the desired product in the balanced equation. $\text{Atom Economy} = \frac{\text{mass of desired products}}{\text{total mass of reactants}} \times 100$		☹	😐	☺
140	The atom economy of a reaction can be calculated using correct formula: e.g. calculate the atom economy of $C_9H_8O_4$ in the following reaction. $C_7H_6O_3 + C_4H_6O_3 \longrightarrow C_9H_8O_4 + C_2H_4O_2$ 1mol 1mol 1mol 1mol 138g 102g 180g $\text{atom economy} = \frac{\text{mass of desired product}}{\text{total mass of reactants}} \times 100 = \frac{180\text{g}}{240\text{g}} \times 100 = 75\%$		☹	😐	☺
141	Different routes are taken in manufacturing products depending on percentage yield and atom economy: • high percentage yield and low atom economy results in waste products being produced • industrial process want efficiency i.e. high percentage yield and high atom economy.		☹	😐	☺

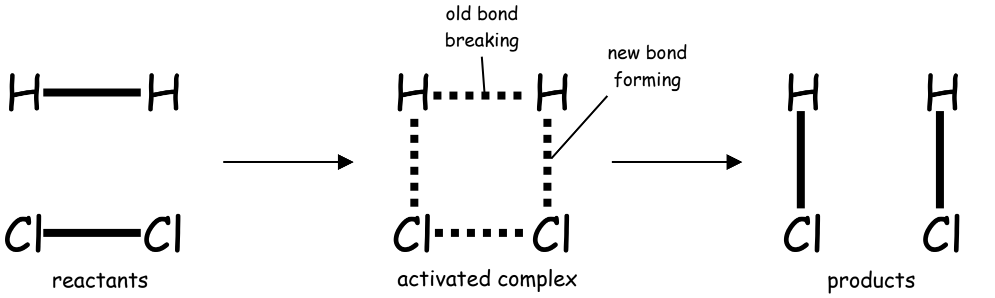
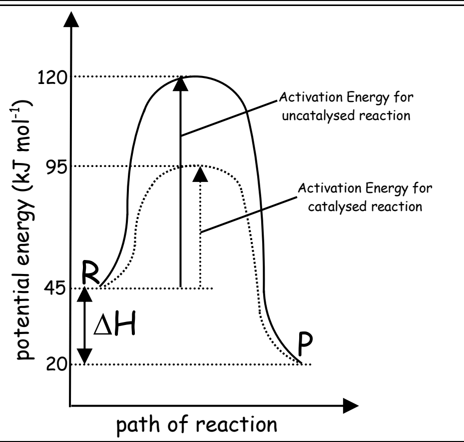
142	Reaction rates must be controlled by in industrial processes. • If the rate is too low the process will not be economically viable • If the rate id too high the process will have a risk of explosion		☹	😐	☺
143 (3)	The relationship between reaction time and reaction rate is: $\text{rate} = \frac{1}{\text{time}}$ • Units of rate include s^{-1}		☹	😐	☺
144 (5)	Collision Theory can be used to explain the effects of concentration, pressure, particle size, temperature on reaction rates using collision geometry a) increased concentration gives a greater chance of a collision ∴ faster reaction b) increased pressure increases chances of a collision ∴ faster reaction c) smaller particles have a larger surface area so more particles available to react by collision d) increased temperature results in more particles moving faster i. faster particles results in more collisions ii. more particles have energy greater than the activation energy ∴ faster reaction e) Collision theory states that before a reaction can take place, the particles must collide with each other with the correct energy and the correct angle of collision for a collision to be a successful collision to form products		☹	😐	☺

Table for 143:

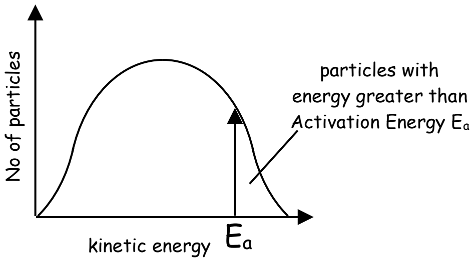
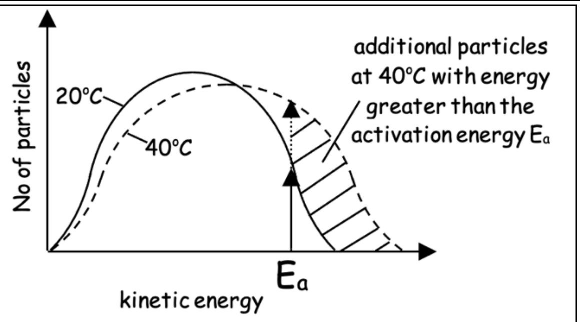
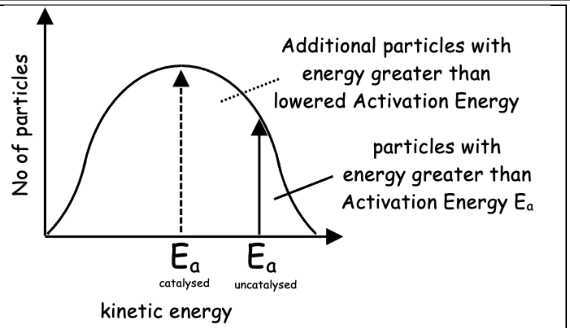
Temperature (°C)	**Time Taken** (s)	**Relative Rate** (s^{-1})
20	100	$^1/_{100}$ = 0.01
30	50	$^1/_{50}$ = 0.02
40	10	$^1/_{10}$ = 0.10

	Higher Chemistry Self-Evaluation Unit 3.2b Reaction Pathways		Page	Traffic Light Red	Amber	Green

145	A potential energy diagram can be used to show the energy pathway for a reaction.
	• The enthalpy change (ΔH) is the energy difference between the products and the reactants
	• Enthalpy change is measured from R to P for the forward reaction.
	• Enthalpy change is measured from P to R for the reverse reaction
	  ○ enthalpy change has a negative value for exothermic reactions (downhill overall)
	  ○ enthalpy change has a positive value for endothermic reactions (uphill overall)
				☹	😐	☺

146a (27)	The activation energy (E_a) is the minimum energy required by colliding molecules to form an activated complex • Activation energy is measured from R to top of hill for the forward reaction • Activation energy is measured from P to top of hill for the reverse reaction • activation energy is the minimum kinetic energy required by colliding molecules for a reaction to occur	<u>For Forward Reaction</u>: E_a = 120 – 45 = 75 kJ mol^{-1} ΔH = 20 – 45 = -25 kJ mol^{-1} <u>For Reverse Reaction</u>: E_a = 120 – 20 = 100 kJ mol^{-1} ΔH = 45 – 20 = +25 kJ mol^{-1}		☹	😐	☺
146b (26)	The activated complex is an unstable arrangement of atoms formed at the maximum of the potential energy barrier (top of hill), during a reaction. 			☹	😐	☺
147 **148**	A catalyst provides an alternate reaction pathway with a lowering of the activation energy • activation energy of reverse reaction is also lowered Enthalpy of reactants and products is unchanged • adding a catalyst has no effect on the enthalpy change (ΔH)			☹	😐	☺

Higher Chemistry Self-Evaluation
Unit 3.2c Kinetic Energy Distribution

		Page	Red	Amber	Green
149	Temperature is a measure of the average kinetic energy of the particles in a substance.		☹	😐	☺
150	The activation energy is the minimum kinetic energy required by colliding particles before a reaction may occur		☹	😐	☺
151 (6) (7) (8) (9)	Energy Distribution Diagrams can be used to explain the effect of changing temperature on the kinetic energy of particles and reaction rate. • Only particles with energy greater than the activation energy can react during a collision 		☹	😐	☺
152a	Energy distribution diagrams can explain the effect of changing temperature on the kinetic energy of particles. • An increase in temperature increases the number of particles with energy greater than the activation energy. • Increase in temperature moves curve to right • Decrease in temperature moves curve to left 		☹	😐	☺
152b	Catalysts lower the activation energy for a reaction • Easier for the activated complex to form as minimum energy required to form activated complex is reduced by adding a catalyst 		☹	😐	☺

		Page	Red	Amber	Green
153	Enthalpy (H) is a measure of the energy stored in a chemical. • Enthalpy change is given the symbol ΔH		☹	😐	☺
154 **155**	A reaction or process that releases heat energy is described as **exothermic** • in industry, exothermic reactions may require heat to be removed to prevent the temperature rising A reaction or process that releases heat energy is described as **endothermic** • in industry, endothermic reactions may incur costs in supplying heat energy in order to maintain the reaction rate.		☹	😐	☺
156 **157** **158** (34)	The enthalpy change for a reaction can be calculated from the data for specific heat capacity, mass and temperature change. • By calculation of how much of 1 mole was burned, the enthalpy of combustion can be calculated e.g. Calculate the enthalpy of combustion of ethanol if 0.92g of ethanol burned to heat up 200cm^3 of water by 6°C. $E_h = c \times m \times \Delta T$ $= 4.18 \times 0.2 \times 6$ $= 5.016$ kJ c = specific heat capacity = 4.18 kJ kg^{-1} °C^{-1} m = mass of water being heated up (worked out by converting volume of water into mass) (NB 1000cm^3 water = 1kg of water) 1mol of ethanol C_2H_5OH = (2x12)+(6x1)+(1x16) = 24 + 6 + 16 = 46g 0.92g ethanol $\longleftrightarrow$ 5.016kJ 46g $\longleftrightarrow$ 5.016kJ x $^{46}/_{0.92}$ = 250.8 kJ mol^{-1} but exothermic reaction ΔH= -250.8 kJ mol^{-1}		☹	😐	☺
159 (31)	The enthalpy of combustion of a substance is the amount of energy given out when one mole of a substance burns completely in oxygen.		☹	😐	☺

Higher Chemistry Self-Evaluation
Unit 3.3b Hess's Law

JAB chem

		Page	Traffic Light Red	Amber	Green
160a	<u>Hess's Law</u>: Enthalpy change for any particular chemical reaction is the same regardless of chemical route taken.		☹	😐	☺
160b	Enthalpy changes can be calculated by application of Hess's Law: e.g. Calculate the enthalpy of formation for SiH_4 $\boxed{Si + 2H_2 \longrightarrow SiH_4}$ ❶ $SiH_4 + 2O_2 \longrightarrow SiO_2 + 2H_2O$ $\Delta H = -1517$ kJ ❷ $Si + O_2 \longrightarrow SiO_2$ $\Delta H = -911$ kJ ❸ $H_2 + \frac{1}{2}O_2 \longrightarrow H_2O$ $\Delta H = -286$ kJ ❶x-1 $SiO_2 + 2H_2O \longrightarrow SiH_4 + 2O_2$ $\Delta H = +1517$ kJ ❷ $Si + O_2 \longrightarrow SiO_2$ $\Delta H = -911$ kJ ❸x2 $2H_2 + O_2 \longrightarrow 2H_2O$ $\Delta H = -572$ kJ add $Si + 2H_2 \longrightarrow SiH_4$ $\Delta H = +34$ kJ		☹	😐	☺
	e.g. calculate the enthalpy of formation of ethyne $\boxed{2C + H_2 \rightarrow C_2H_2}$ ❶ $C + O_2 \rightarrow CO_2$ $\Delta H = -394\text{kJ mol}^{-1}$ ❷ $H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ $\Delta H = -286\text{kJ mol}^{-1}$ ❸ $C_2H_2 + 2\frac{1}{2}O_2 \rightarrow 2CO_2 + H_2O$ $\Delta H = -1300\text{kJ mol}^{-1}$ ❶x2 $2C + 2O_2 \rightarrow 2CO_2$ $\Delta H = -788\text{kJ mol}^{-1}$ ❷ $H_2 + \frac{1}{2}O_2 \rightarrow H_2O$ $\Delta H = -286\text{kJ mol}^{-1}$ ❸x-1 $2CO_2 + H_2O \rightarrow C_2H_2 + 2\frac{1}{2}O_2$ $\Delta H = +1300\text{kJ mol}^{-1}$ add $2C + H_2 \rightarrow C_2H_2$ $\Delta H = +226\text{kJ mol}^{-1}$		☹	😐	☺

Higher Chemistry Self-Evaluation
Unit 3.3c Bond Enthalpies

JAB chem

		Page	Traffic Light Red	Amber	Green
161	Molar Bond Enthalpy is the energy required to break one mole of bonds in a doatomic molecule. • Mean molar bond enthalpy is the average energy required to break one mole of bonds, for a bond that occurs in a number of compounds. e.g. 1mol of C-H bonds requires 412 kJ of energy to break 1mol of C-H bonds releases 412 kJ of energy when formed		☹	😐	☺
162	Bond enthalpies can be used to calculate the enthalpy change for reactions in the gas phase: e.g. calculate the enthalpy of formation of HCl: $\frac{1}{2}H_{2(g)} + \frac{1}{2}Cl_{2(g)} \longrightarrow HCl_{(g)}$ <u>Endothermic Steps: Bond Breaking</u> $\frac{1}{2}$mol H-H $\frac{1}{2} \times +436\text{kJ} = 218.0\text{kJ}$ $\frac{1}{2}$mol Cl-Cl $\frac{1}{2} \times +243\text{kJ} = 121.5\text{kJ}$ 339.5kJ <u>Exothermic steps: Bond forming</u> 1mol H-Cl 432kJ 432kJ ∴ Enthalpy Change = (Total of Bond Breaking Steps) – (Total of Bond Forming Steps) = 337.5kJ – 432kJ = -94.5kJ mol^{-1}		☹	😐	☺

Higher Chemistry Self-Evaluation
Unit 3.4 Equilibria

		Page	Red	Amber	Green
163 (16)	Reversible reactions attain a state of dynamic equilibrium when the rates of forward and reverse reactions are equal • reversible reactions are reactions where the forward reaction and the reverse reactions both take place at the same time • Le Chatelier's Principle states: An equilibrium will move to undo any change imposed upon it by temporarily favouring either the forward or backward reaction until equilibrium is reached again.		☹	😐	☺
164 (17)	At equilibrium, the concentrations of reactants and products remain **constant**, • concentrations of reactants and products are unlikely to be equal at equilibrium • the reaction has not stopped at equilibrium		☹	😐	☺
165	The chemical industry employs strategies to move equilibrium in favour of making more products		☹	😐	☺
166a (18c)	Le Chatelier's Principle can explain the effect on the equilibrium position of changing temperature (see below)		☹	😐	☺
166b (18b)	Le Chatelier's Principle can explain the effect on the equilibrium position of changing pressure (see below)		☹	😐	☺
166c (18a)	Le Chatelier's Principle can explain the effect on the equilibrium position of changing concentration (see below)		☹	😐	☺
167 (19)	A catalyst increases the rate of the forward reaction and reverse reaction equally. • catalyst increases rate at which equilibrium is achieved but no change to position of equilibrium • final percentage of products is the same with or without a catalyst at equilibrium		☹	😐	☺
- (20)	The Haber Process is a very important industrial process (see below)		☹	😐	☺

166a (18c) Le Chatelier's Principle can explain the effect on the equilibrium position of changing temperature:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.4\,kJ\,mol^{-1}$$

Forward Reaction is **exo**thermic — Reverse Reaction is **endo**thermic

Increase in Temperature	Decrease in Temperature
System fights back by trying to lower temperature	System fights back by trying to raise temperature
Temperature reducing reaction is favoured ∴ endothermic reaction is favoured	Temperature increasing reaction is favoured ∴ exothermic reaction is favoured
Equilibrium moves to the left ∴ *less* products at equilibrium	Equilibrium moves to the right ∴ *more* products at equilibrium

166b (18b) Le Chatelier's Principle can explain the effect on the equilibrium position of changing pressure:

$$\underbrace{\underset{\substack{1mol\\\downarrow\\1vol}}{N_2(g)} + \underset{\substack{3mol\\\downarrow\\3vol}}{3H_2(g)}}_{} \rightleftharpoons \underbrace{\underset{\substack{2mol\\\downarrow\\2vol}}{2NH_3(g)}}_{}$$

$$4vol\ of\ gas \rightleftharpoons 2vol\ of\ gas$$

Increase in Pressure	Decrease in Pressure
System fights back by trying to lower pressure	System fights back by trying to raise pressure
Pressure reducing reaction is favoured ∴ forward reaction is favoured	Pressure increasing reaction is favoured ∴ reverse reaction is favoured
Equilibrium moves to the right ∴ *more* products at equilibrium	Equilibrium moves to the left ∴ *less* products at equilibrium

166c (18a) Le Chatelier's Principle can explain the effect on the equilibrium position of changing concentration:

Change	Effect on Equilibrium		
Addition of a reactant	Equilibrium tries to remove additional reactant	Forward reaction favoured	Equilibrium shifts to right
Removal of a reactant	Equilibrium tries to replace removed reactant	Reverse reaction favoured	Equilibrium shifts to left
Addition of a product	Equilibrium tries to remove additional product	Reverse reaction favoured	Equilibrium shifts to left
Removal of a product	Equilibrium tries to replace removed product	Forward reaction favoured	Equilibrium shifts to right

- (20) The Haber Process is a very important industrial process:

a) the effects of pressure
- high pressure increases the products at equilibrium

b) use of temperature
- moderate temperatures are used as high temperatures favour the reverse reaction which breaks down ammonia back to the reactants

c) the use of a catalyst
- use of an iron catalyst increases the rate of ammonia production by achieving equilibrium more quickly (but does not produce more ammonia at equilibrium)

d) recycling of unreacted gases
- ammonia is easily separated from unreacted nitrogen & hydrogen as ammonia has a much higher boiling point. Unreacted nitrogen & hydrogen are returned to the reaction vessel

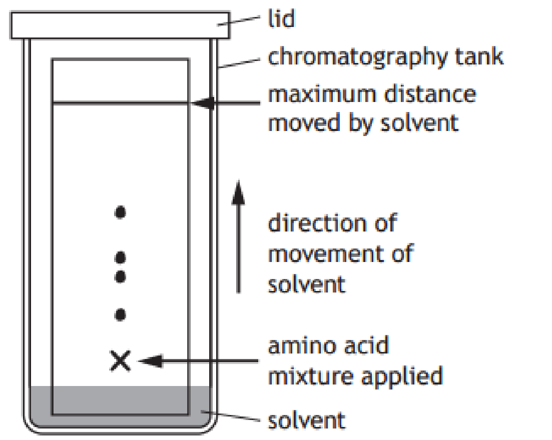
e) removal of product
- removal of ammonia product before equilibrium is achieved means the system tried to replace ammonia to try to achieve equilibrium

		Page	Traffic Light: Red	Amber	Green

168	Chromatography is a technique to separate the components present within a mixture.
	• chromatography separates substances by making use of differences in their polarity or molecular size
	• the relative affinity for the mobile phase and the stationary phase decides how far/fast the substances travels during chromatography
	• mobile phase is a liquid or a gas which carries the sample through the material.
	  ○ size of molecules and their polarity affect how soluble they are in the mobile phase (and how far they travel in the mobile phase)
	• stationary phase may be paper, silica gel, or an inert packing material in a column.
	  ○ size and polarity of the compounds may affect their affinity for the stationary phase (how little they travel in the mobile phase)
			☹	😐	☺

169a	Chromatography can identify a component by the distance the component travelled. (see below)				
169b 170	Chromatography can be used to identify a component by the time taken the component to travel through the apparatus (known as the retention time). • The quantity of the substance is indicated by the height of the peak • The retention time of substance indicated the size and polarity of the substance		☹	😐	☺

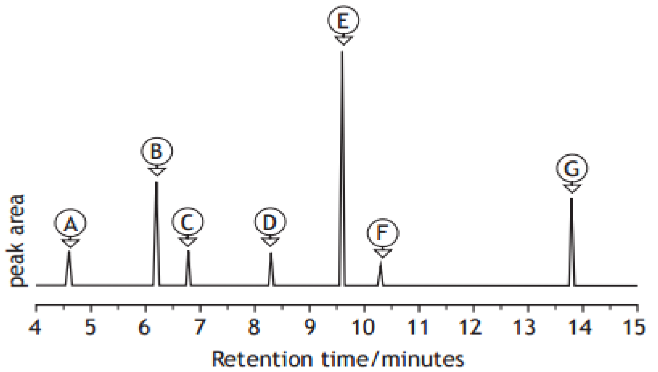
169a



Each chemical in the sample travels a different difference and can be quantified by the R_f value:

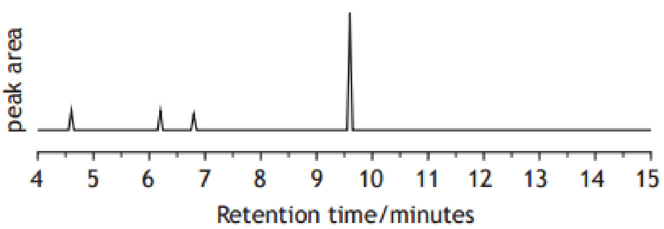
$$R_f = \frac{\text{distance travelled by sample}}{\text{distance travelled by solvent}}$$

169b / 170



The brand name perfume gives the following gas chromatogram showing varying quantities of 7 different chemicals

- (A) linalool
- (B) citronellol
- (C) geraniol
- (D) eugenol
- (E) anisyl alcohol
- (F) coumarin
- (G) benzyl salicylate



The counterfeit brand of perfume contains some but not all peak of the brands name perfume:

Retention Time (min)	Chemical Identified
4.6	A (linalool)
6.2	B (citronellol)
6.8	C (geraniol)
9.6	E (anisyl alcohol)

171	Volumetric analysis uses a solution of accurately known concentration to determine the exact concentration of another substance using titration. • an exact volume and concentration of a substance will allow the calculation of the number of moles of a substance. • using the mole ratio from a balanced equation, the number of moles of a second substance can be calculated • the exact volume of the second substance, measured accurately using a burette, will allow the calculation of the concentration of the second substance.		☹	😐	☺
172	Titration is used to accurately determine the volume of solution to reach the end-point of a reaction. • indicator is used to show when the end-point has been reached. • titre volumes with 0.2cm^3 are considered concordant with the rough titration ignored.		☹	😐	☺
173	Standard solutions are solutions with an accurately known concentration. • dissolve a accurately measured mass of solid in a small volume of deionised water in a beaker • transfer the solution to a standard flask, rinsing the beaker carefully • make up the solution to the mark on the standard flask, using a dropper for the last few drops so that the bottom of the meniscus is touching the line on the flask.		☹	😐	☺
174	Redox titrations are based on redox reactions. • Titrations using acidified permanganate solution are self-indicating as purple permanganate turns colourless as the permanganate ions are reduced.		☹	😐	☺
175	Titrations and Balanced Redox Equations are used to calculate the concentration of a reactant, given the concentration of the other. (see below)		☹	😐	☺

Titrations and Balanced Redox Equations are used to calculate the concentration of a reactant, given the concentration of the other.

Question

The vitamin C content in a fruit drink can be determined by titrating it with iodine.

$$\underset{\text{Vitamin C}}{C_6H_8O_6(aq)} + I_2(aq) \longrightarrow C_6H_6O_6(aq) + 2H^+(aq) + 2I^-(aq)$$

To determine the vitamin C content in a 1·0 litre carton of orange juice, three separate 20cm^3 samples of the juice were titrated with a 0·00125mol l^{-1} iodine solution. Starch indicator was used to determine the endpoint with a colourless to blue/black colour change.

The following results were obtained from titration of the three 20cm^3 samples of orange juice.

Titration	Volume of 0.00125mol l^{-1} iodine solution used /cm^3
1	26.3
2	25.5
3	25.3

Calculate the concentration, in mol l^{-1}, of vitamin C, in the 1·0 litre carton of orange juice.

Solution to Problem

$$\text{Average titre} = \frac{25.3+25.5}{2} = \frac{50.8}{2} = 25.4\text{cm}^3$$

no. of mol I_2 = volume x concentration = 0.0254litres x 0.00125mol l^{-1} = 3.175x10^{-5}mol

$C_6H_8O_6$	+	I_2	$\longrightarrow$	$C_6H_6O_6$	+	$2H^+$	+	$2I^-$
1mol		1mol						
3.175x10^{-5}mol		3.175x10^{-5}mol						

∴ 20cm^3 orange juice contains 3.175x10^{-5}mol Vitamin C ($C_6H_8O_6$)

$$\text{concentration} = \frac{\text{no. of mol}}{\text{volume}} = \frac{3.175\times10^{-5}\text{ mol}}{0.020\text{ litres}} = 0.00159\text{ mol l}^{-1}$$

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