Year 12 Revision Questions –
Unit 3 Analytical & Equilibrium Chemistry

Q1.

Which of the following represent example(s) of quantitative analysis?

I A sample of 'fries seasoning' contains salt, chicken extract, rice flour, paprika and chilli.
II Dry air contains 78% nitrogen vol/vol.
III A sample of iron ore weighs 23.5 kg.
IV A chocolate biscuit contains 2.3 g protein, 6.6 g fat and 24.4 g carbohydrate per 100 g serve.

A IV only.
B I and IV.
C II and IV.
D II, III and IV.

Q2.

A chemist used atomic emission spectroscopy to confirm the presence of calcium in a sample of soft drink. Atomic emission spectroscopy was used in preference to a flame test because:
A atomic emission is cheaper than a flame test.
B atomic emission is quicker than a flame test.
C calcium salts do not produce a colour in a flame test.
D atomic emission can distinguish between metals that produce similar colours in a flame test.

Q3.

Which of the following techniques would be best to analyse for the concentration of aspirin in a headache remedy?
A colorimetry
B paper chromatography
C high performance liquid chromatography
D atomic absorption spectroscopy

Q4.

Which of the following techniques would be best to analyse for low concentrations of lead in blood?
A gas–liquid chromatography

B atomic absorption spectroscopy
C flame test
D electrophoresis

Q5.

Sodium salts give a distinctive yellow-orange colour in a flame test. The colour is due to:
A electrons near the nucleus of the sodium atom absorbing energy.
B excited electrons returning to lower energy levels and releasing energy.
C sodium atoms losing an electron to form sodium ions.
D sodium ions being converted to sodium atoms in the reducing flame.

Q6.

Three of the components of the green colouring in grass and their R_f values are β-carotene 0.82, chlorophyll *a* 0.48 and chlorophyll *b* 0.35. The green pigment is analysed by paper chromatography and the solvent front travels 10 cm. Which of the following statements is correct?
A Chlorophyll *b* is the most strongly adsorbed component.
B The band due to β-carotene is nearest the origin.
C The band due to chlorophyll *a* will be 0.48 cm from the origin.
D The band due to β-carotene is 0.82 cm from the solvent front.

Q7.

Which of the following are examples of concentration units?
I mg/kg
II parts per million (ppm)
III tonne
IV percentage (%)

A I only.
B II only.
C II and III.
D I, II and IV.

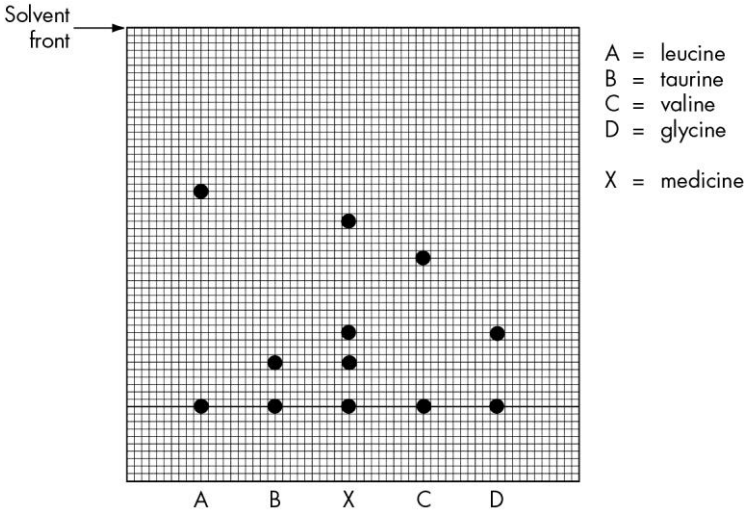
Q8.

Which of the following is/are a feature of both atomic emission spectroscopy and flame tests?
I Use of a very hot ethyne flame.
II Excitation of metal atoms.
III Passage of light through a prism.
I VSeparation of the wavelengths of light into a spectrum.

A I only.

B II only.
C II and III.
D III and IV.

Q9.



a The following materials are commonly used in chromatography: water, ethanol, paraffin wax, glass sheets, paper strips, powdered alumina, nitrogen gas and hexane.
Choose a suitable stationary phase and mobile phase from the list to use in the analysis of amino acids by thin layer chromatography.

b Calculate the distance moved by the sample of leucine if the solvent front moves 13 cm from the origin.

c The separation of taurine and glycine on the chromatogram is not very great. State what would happen to the R_f values of the two amino acids if the separation was carried out for a longer period so the solvent front moved 10 cm instead of 5 cm from the origin. Explain your answer.

d Write a short paragraph explaining how leucine is separated from the other amino acids.

Q10.

a A sample of common salt, sodium chloride, gives a yellow colour in the gas flame when spilt on the stove. Use a labelled diagram to assist your explanation of how the coloured light is formed.

b Calcium forms a red colour in a flame test while copper turns the flame blue/green. What can you deduce about the most frequent electronic transition that occurred in each case?

c The concentration of calcium in blood was determined by atomic emission. Give two reasons why a flame test is not used.

Q11.

a Distinguish between paper chromatography and atomic emission spectroscopy in terms of the following features.
   i The type of sample analysed, for example metals, inorganic compounds, organic compound, anions, cations etc.
   ii Whether atoms or compounds are determined.
   iii Whether separation of the components of the sample occurs.
   iv Quantitative or qualitative analysis.
   v Cost of the test.

b Paper chromatography and atomic emission spectroscopy involve very different chemical principles. Below are four key terms that can be used in an explanation of each of the techniques. For each of the four terms identify whether it applies to paper chromatography or atomic emission spectroscopy and use the key term in a sentence to explain its meaning.
   i desorption
   ii energy levels
   iii excited electron
   iv solvent front

Q12.

A sample of a gaseous hydrocarbon, molecular formula C_2H_4, was heated in excess oxygen and then cooled to room temperature. The initial volume of the hydrocarbon, measured under standard laboratory conditions (SLC) was 10 L. Assuming complete combustion, what was the total volume of the gaseous products after cooling to room temperature?

A 5.0 L
B 10 L
C 20 L
D 40 L

Q13.

A number of ammonia-based compounds are used as nitrogen fertilisers. Which of the following has the greatest percentage by mass of N?

A ammonium nitrate
B ammonium sulfate
C ammonium phosphate
D urea $(NH_2)_2CO$

Q14.

A 10 g sample of blue hydrated copper sulfate is heated in a crucible over a Bunsen flame until the crystals have turned white. The mass of the white residue is found to be 6.9 g. Assume that the mass lost is due to loss of the water of hydration.

On the basis of these data, the empirical formula of the hydrated copper sulfate is:

A $CuSO_4.7H_2O$
B $CuSO_4.6H_2O$
C $CuSO_4.4H_2O$
D $CuSO_4.3H_2O$

Q15.

The concentration of sodium chloride in seawater can be determined by a number of analytical techniques. Which of the following methods can be used?

I gravimetric analysis
II flame tests
III atomic emission spectroscopy
IV thin layer chromatography

A I and II.
B I and III.
C I, II and III.
D I, II, III and IV.

Q16.

Which of the following properties is an important property of a precipitate used in gravimetric analysis? The precipitate should:

A be stable to heat.
B be coloured.
C have a low molar mass.

D form solids with most ions.

Q17.

Dried rice is analysed to make sure that it contains only a very small percentage of water. The rice sample is repeatedly weighed, heated, cooled and reweighed until a constant mass is obtained. The sample is cooled in a desiccator before weighing.

A desiccator is used to ensure that the sample:

A will not absorb water from the atmosphere.
B will not absorb oxygen from the atmosphere.
C will not release water to the atmosphere.
D will not release volatile component to the atmosphere.

Q18.

Xylene, a colourless organic liquid, is used as an industrial solvent. A 9.0 g sample of xylene, equivalent to 0.085 mol, was found to contain 90.6% carbon and 9.4% hydrogen. Determine the empirical and molecular formula of xylene.

Q19.

Liquid ammonia, NH_3, is widely used by farmers as a fertiliser. Industrially, ammonia is produced by the Haber process. Hydrogen and nitrogen gases are introduced into a steel container, in the stoichiometric proportions of the equation. A total mass of 340 g gases are introduced into a 20 L steel container at 25°C.

a Write a balanced equation for the production of ammonia from its elements.
b The total mass of hydrogen and nitrogen is 340 g. What is the mass of each gas if mixed in stoichiometric proportions?
c Calculate the moles of hydrogen and nitrogen and thus the total moles of gas at the start of the reaction.
d What is the total pressure in the container due to hydrogen and nitrogen before the reaction starts? Give your answer in atmospheres.

Q20.

In black and white films, unexposed silver bromide is removed by reaction with sodium thiosulfate solution:

$$AgBr(s) + 2Na_2S_2O_3(aq) \rightarrow Na_3Ag(S_2O_3)_2(aq) + NaBr(aq)$$

a 1.5 g of unexposed silver bromide is present on the film. The film is treated with a 0.15 M solution containing 2.0 g of sodium thiosulfate.
   i Which reactant is in excess?
   ii What mass of sodium bromide is formed?
   iii What is the minimum volume of sodium thiosulfate needed to just react with the unexposed silver bromide.

b Aqueous silver ions from used photographic solutions can be recovered by precipitation as silver chloride. Silver chloride has a very low solubility in water. However, a very small amount will dissolve, causing an error in the calculated mass of silver obtained.
   i Write an ionic equation for the formation of silver chloride precipitate from aqueous silver ions.
   ii Will the slight solubility of silver chloride result in a value that is higher or lower than the true value. Explain your answer.
   iii Is the difference from the true value in part b ii due to a systematic or random error?
   iv Give an example of the other type of error.

Q21.

Which list best represents the relative concentrations of some of the species present in a 1 M aqueous solution of ammonia?

A $[H_2O] > [NH_3] > [OH^-] > [H_3O^+]$
B $[NH_3] > [NH_4^+] > [OH^-] > [H_2O]$
C $[NH_4^+] = [OH^-] > [H_2O] > [NH_3]$
D $[NH_3] = [NH_4^+] = [OH^-] > [H_3O^+]$

Q22.

Which of the following is *not* used as a primary standard in volumetric analysis?

A hydrated oxalic acid
B sodium hydroxide
C sodium oxalate
D anhydrous sodium carbonate

Q23.

Which of the following statements about acid–base indicators is *incorrect?*

A Acid–base indicators are often weak acids.
B Acid–base indicators have an acid form that has a different colour from its conjugate base.
C Acid–base indicators are present only in the acid form at the equivalence point of the titration.
D Acid–base indicators can be extracted from a number of plant materials.

Q24.

A standard 0.103 M sodium carbonate solution was prepared in a volumetric flask. Aliquots of 25.00 mL were measured into a conical flask and titrated against a dilute solution of approximately 0.2 M hydrochloric acid. The following titres were obtained: 22.35, 22.45, 25.05 and 22.35 mL.

An average titre is calculated from concordant results. The best value for the average titre is:

A 20.00 mL
B 22.35 mL
C 22.38 mL
D 23.05 mL

Q25.

A standard 0.103 M sodium carbonate solution was prepared in a volumetric flask. Aliquots of 25.00 mL were measured into a conical flask and titrated against a dilute solution of approximately 0.2 M hydrochloric acid. The following titres were obtained: 22.35, 22.45, 25.05 and 22.35 mL.

One group of students found that their calculated value of hydrochloric acid was lower than that of all their classmates. Which of the following errors would account for this error?

A Using a 20.00 mL pipette to measure the base but basing the calculation on a 25.00 mL pipette.
B Rinsing the burette with water instead of acid.
C Adding 8 drops of indicator rather than four.
D Using an impure sample of sodium carbonate to prepare the sodium carbonate solution.

Q26.

A mixture of 10 g of sodium carbonate and 10 g of sodium hydroxide was mixed and then dissolved in water. What volume of 0.10 M hydrochloric acid is needed to neutralise the solution?

A 0.9 L
B 1.9 L
C 2.5 L
D 4.4 L

Q27.

An open-air swimming pool contains 120 000 L of salt water. After heavy winter rain the salt content was found to be 25 g L^{-1}. If 720 kg sodium chloride were added what would be the resulting molarity of the solution?

A 0.10 M
B 0.43 M
C 0.49 M
D 0.53 M

Q28.

The drain cleaner 'Drainol' contains 13% m/v NaOH. The original solution is diluted, and 10 mL is made up to 500 mL with distilled water.

a What is the molarity of the original solution?
b Concentrated solutions of NaOH are not used in a titration. Give two reasons why.
c The concentration of the diluted sodium hydroxide is determined by titration. Give a suitable reactant and indicator to use in the titration with sodium hydroxide.
d Standard sodium hydroxide solutions cannot be prepared directly from the solid. Explain why.

Q29.

A student wishes to prepare a standard solution of a base for an acid–base titration. She chooses to weigh the required amount of anhydrous sodium carbonate, dissolve the solid in water and make up the volume in a volumetric flask. The 0.0500 M standard sodium carbonate solution is then used to determine the amount of hydrochloric in a sample of concrete cleaner. 20.00 mL of the diluted concrete cleaner is pipetted into the conical flask.

After several titrations a 23.45 mL titre is obtained and the concentration of the hydrochloric acid is calculated.

a Distinguish between the terms primary standard and standard solution.

b List three features that make anhydrous sodium carbonate a suitable choice of solid to use to prepare the standard solution.

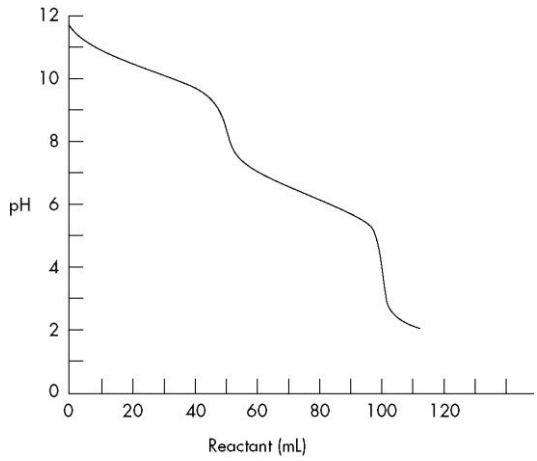
c Which is the best answer for the concentration of hydrochloric acid in the diluted concrete cleaner: 0.1 M, 0.11 M, 0.117 M, 0.1172 M, 0.1173 M, 0.11725 M?

d Determine the effect of the following errors on the final calculated value of the hydrochloric acid concentration.

i The burette was rinsed with water.

ii The conical flask was rinsed hydrochloric acid.

Q30.



The titration curve of this acid–base reaction was measured by taking continuous readings from a pH electrode in the conical flask. Solution A (0.1 M, 5 0 mL) was placed in a conical flask and titrated against 0.1 M solution B.

Possible reactants and indicators are hydrochloric acid, acetic (ethanoic) acid, ammonia solution, sodium carbonate solution, sodium hydroxide, methyl orange (end point pH 3.2–4.4), phenolphthalein (end point pH 8.2–10.0).

a Which reactant in the conical flask could account for the titration curve shown?

b Give the pH values at which the equivalent points occur.

c Suggest an appropriate indicator to determine the first equivalence point.

d Write a molecular equation, using appropriate reagents, describing the reaction that is represented by the second equivalence point.

e Distinguish between the terms equivalence point and end point in a titration.

Q31.

The following paragraph describes an acid–base titration. Some of the key words are missing. Use the list of words in the box to fill in the gaps.

pipette measuring cylinder beaker volumetric flask primary standard indicator base standard solution burette indicator titre desiccator aliquot

An approximate 2 g sample of anhydrous sodium carbonate is accurately weighed. (The solid must be dry if it is to be used as a _____.) The solid is tipped into a _____ and shaken with about 50 mL of distilled water until the solid dissolves. More water is added to make the solution to a volume of exactly 100.0 mL. A 20.00 mL _____ of the solution is taken by using a _____ and placed in a conical flask. A few drops of methyl orange are added and the mixture is titrated against dilute hydrochloric acid.

Q32.

An impure sample of limestone, mainly calcium carbonate, was analysed by using a back titration. Approximately 1 g of the finely powdered limestone was accurately weighed into a conical flask. An excess of the hydrochloric acid, exactly 50.00 mL, was added to the limestone. The mixture was stirred for 15 min with a magnetic stirrer to allow the reaction to be completed. The hydrochloric acid was titrated with a standard solution of sodium hydroxide.

The following results were obtained:

- Mass of watch glass = 8.7954 g
- Mass of anhydrous sodium carbonate and watch glass = 9.8460 g
- Concentration of standard sodium hydroxide solution = 0.0489 M
- Titration value of sodium hydroxide obtained = 22.32 mL
- Concentration of hydrochloric acid = 0.395 M

a Write balance equations for the two reactions that occur.

b Determine the moles of hydrochloric acid in excess after the reaction with the limestone.

c Calculate the total moles of hydrochloric acid added to the limestone.

d How many moles of hydrochloric acid reacted with the limestone?

e Calculate the number of moles of calcium carbonate in the limestone.

f What is the percentage of calcium carbonate in the limestone.

g In this experiment the whole sample of limestone was used in one titration. How could the precision of the titration have been improved?

Q33.

Consider the equation:

$$H_2C_2O_4(aq) + MnO_4^-(aq) + H^+(aq) \rightarrow CO_2(g) + Mn^{2+}(aq) + H_2O(l)$$

This is an unbalanced equation. When correctly balanced, the mole ratios of the products (CO_2 + Mn^{2+} + H_2O) are:

A 2 : 1 : 3
B 5 : 1 : 8
C 5 : 2 : 6
D 10 : 2 : 8

Q34.

The oxidation number of C in $H_2C_2O_4$ is:
A 0
B +2
C +3
D +4

Q35.

The oxidation number of Br in BrO_3^- is:
A –1
B 0
C +5
D +7

Q36.

In a titration to determine the concentration of iodine in an antiseptic solution, the following reaction occurred:

$$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$$

Starch indicator is added to the thiosulfate, $S_2O_3^{2-}$, in the conical flask and titrated against I_2 until a permanent blue colour is observed.

Which of the following statements about the end point/equivalence point is true?
A The end point is reached before the equivalence point.
B The end point is reached after the equivalence point.
C The end point is reached at the same time as the equivalence point.
D The amount of starch added determines whether the end point or equivalence point is reached first.

Q37.

What volume of 2.0 M stock sodium iodide solution is needed to prepare 500 mL of 0.3 M sodium iodide solution?
A 25 mL
B 75 mL
C 150 mL
D 300 mL

Q38.

Consider the equation:

$$3N_2H_5^+(aq) + 4NO_3^-(aq) + H^+(aq) \rightarrow 3N_2(g) + 4NO(g) + 8H_2O(l)$$

The correct order of increasing oxidation number of N in these compounds is:

A $N_2H_5^+ > N_2 > NO > NO_3^-$

B $N_2 > NO > N_2H_5^+ > NO_3^-$

C $N_2 > NO > NO_3^- > N_2H_5^+$

D $NO > NO_3^- > N_2 > N_2H_5^+$

Q39.

Define the following terms.

a limiting reagent

b standard solution

c average titre

Q40.

A standard solution of sodium oxalate, $Na_2C_2O_4$, was prepared for a redox titration. An accurately known solution of approximately 0.05 M sodium oxalate was required for the titration.

a What mass of sodium oxalate would give a 0.0500 M solution if a 250 mL volumetric flask was used for the dilution?

b Explain why the accuracy of the titration is not affected if the exact mass calculated in part a is not used.

Q41.

'Iron' tablets are commonly prescribed to women in pregnancy. A 200 mg tablet is claimed to contain 50 mg Fe^{2+} (present as iron(II) sulfate). The chemist determined the concentration of iron in the tablet by titrating the dissolved tablet with an acidified solution of potassium permanganate according to the following equation:

$$5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

The tablet was dissolved in approximately 50 mL distilled water. The solution was titrated with a 0.0150 M solution of potassium permanganate. If the calculated concentration of Fe^{2+} was found to be 48.5 mg, what was the titre of potassium permanganate used?

Q42.

A vitamin C (ascorbic acid, $C_6H_8O_6$) tablet, total mass 0.537 g, was analysed by titration to determine the amount of ascorbic acid it contained. The tablet was dissolved in 50 mL of

water. An excess of iodine, 50 mL of 0.0490 M I_2 solution, was added to the dissolved tablet. The excess iodine was titrated with 0.0213 M thiosulfate. The titre recorded was 28.45 mL.

$$C_6H_8O_6(aq) + I_2(aq) \rightarrow C_6H_6O_6(aq) + 2H^+(aq) + 2I^-(aq)$$
$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

a Why is a back titration sometimes used instead of a direct titration?
b Calculate the number of moles of vitamin C in the tablet.
c What is the % w/w vitamin C in tablet?
d If the volume of the water used to dissolve the tablet was less than 50 mL, what effect would there be on the calculated concentration of vitamin C?
e If the volume of excess iodine added to the dissolved tablet was less than 50 mL, what effect would there be on the calculated concentration of vitamin C?

Q43.

The concentration of a substance is most accurately determined in HPLC by measuring:
A peak area
B peak height
C retention time
D R_f value

Q44.

The technique of colorimetry is based on the theory that:
A emission of light is directly proportional to the concentration of the analyte in the cell.
B emission of light is directly proportional to the intensity of the coloured solution in the cell.
C absorption of light is directly proportional to the concentration of the coloured solution in the cell.
D absorption of light is directly proportional to the amount of analyte in the cell.

Q45.

The function of the flame in atomic absorption spectroscopy is to:
I evaporate the solvent.
II convert metal ions to atoms.
III excite the metal atoms.

A I
B I and II.
C I, II and III.
D II and III.

Q47.

The stationary phase in chromatography can be:

I solid
II liquid
III gas

A I
B II
C I or II.
D I, II or III.

Q48.

The concentration of vitamin A is best analysed by:

A gas–liquid chromatography.
B high performance liquid chromatography.
C colorimetry.
D atomic absorption spectroscopy.

Q49.

A sample was found to contain 44 ppm Cu by atomic absorption. This is equivalent to:

A 4.4 μg mL^{-1} Cu.
B 44 μg mL^{-1} Cu.
C 0.044 g L^{-1} Cu.
D 0.44 M Cu.

Q50.

Which of the following techniques require the use of standard solutions to determine the concentration of the analyte (chemical being determined)?

A colorimetry
B gas chromatography
C atomic absorption spectroscopy
D all of the above

Q51.

The function of the monochromator in AAS is to:

A separate the unwanted wavelengths of light from the hollow cathode lamp.

B separate white light from the lamp into its component wavelengths.
C detect the light emitted by the hollow cathode lamp.
D detect the light emitted by the flame.

Q52.

All forms of chromatography involve:
A separation of the components of the mixture.
B a solid stationary phase.
C a solvent that carries the compound along.
D a detector.

Q53.

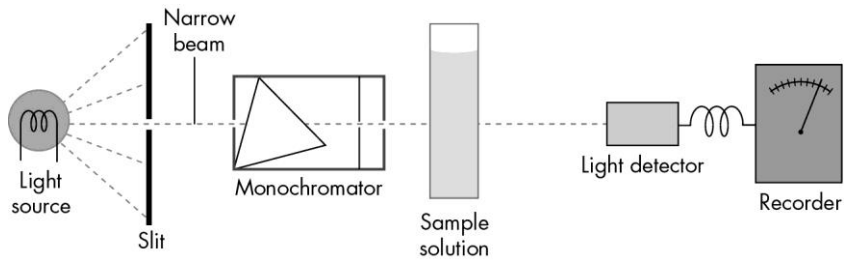
The following forms of spectroscopy—colorimetry, UV, atomic emission and atomic absorption—involve:
A a prism or grating to separate the light.
B a light that is passed through the sample.
C emission of light by the sample.
D a flame to excite the atoms.

Q54.

The most accurate way to measure 5 mL water is by using:
A a burette.
B a pipette.
C an analytical mass balance.
D a volumetric flask.

Q55.

The schematic diagram shows a simple UV–visible spectrophotometer.



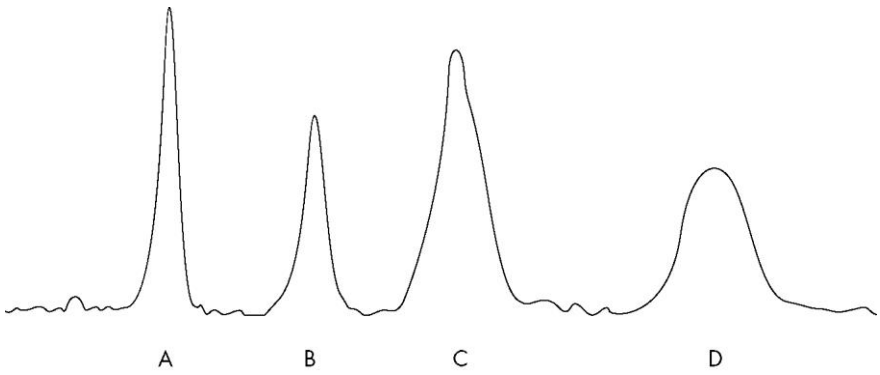
The table shows the colours of visible light and complementary colours.

Wavelength (nm)	Colour absorbed	Colour observed
380–420	Violet	Green-yellow
420–440	Violet-blue	Yellow
440–470	Blue	Orange
470–500	Blue-green	Red
500–520	Green	Purple
520–550	Yellow-green	Violet
550–580	Yellow	Violet-blue
580–620	Orange	Blue
620–680	Red	Blue-green
680–780	Purple	Green

a Name a substance that can be determined directly by UV–visible spectrometry, other than copper sulfate solution.
b Describe the purpose of the monochromator in the machine.
c What wavelength of light would be best used to analyse a sample of copper sulfate solution?
d Describe the changes on the atomic level that occur during UV spectroscopy.

Q56.

A mixture of four alkanes—decane, heptane, hexane and octane—was separated by GLC. The solubility of the hydrocarbons in the liquid phase on the column was directly proportional to the relative molecular mass of the alkane. The chromatogram of the four hydrocarbons is shown.



a Identity the hydrocarbons A, B, C and D.
b Why does decreased solubility of a chemical in the stationary phase decrease the retention time?
c Why do the peaks become broader with increased retention time?
d List the factors that can increase the retention time of a particular chemical in GLC.

Q57.

Various forms of heart disease including angina are very successfully treated with nitroglycerine patches. The small colourless plastic patches are stuck to the skin and slowly release nitroglycerine into the blood at a rate of 0.4 mg/h. The amount of nitroglycerine in the patches must be carefully controlled to prevent an overdose.

A nitroglycerine patch was dissolved in approximately 20 mL ethanol and then diluted to 100 mL. 10.0 mL of this solution was diluted again to 100 mL with water. 20 μL volumes of the diluted sample and 20 μL of each of the prepared standards of nitroglycerine were injected onto a 1.5 m GLC column and analysed by using a nitrogen-specific detector.

Nitroglycerine standards	Peak area (mm^2)
Standard 5 μg/mL	7.2
Standard 10 μg/mL	14.6
Standard 15 μg/mL	22.0
Diluted sample solution	10.8

a Construct a calibration curve and determine the concentration of nitroglycerine in the diluted sample.

b What mass of nitroglycerine was in the patch?

c The patch has a total mass of 0.50 g. What is the percentage concentration (w/w) of nitroglycerine in the patch?

Q58.

Low concentrations of gold can be economically extracted by dissolving the tiny gold particles in the rock with sodium cyanide.

An equation for the reaction is:

$$4Au(s) + 8CN^-(aq) + 2H_2O(l) + O_2(g) \rightarrow 4Au(CN)_2^-(aq) + 4OH^-(aq)$$

a Is this an acid–base or redox reaction? Give an example of an acid/conjugate base or an oxidant/conjugate reductant redox pair in this equation to support your answer.

b What mass of gold could be theoretically extracted with 1.00 L of 0.540 M sodium cyanide solution?

c What technique can be used to analyse gold solutions in concentration of mg Au/L?

Q59.

Conde's crystals, a traditional 'bush' remedy composed of potassium permanganate, can act as a fungicide. The amount of potassium permanganate in a sample of Conde's crystals can be determined by a number of analytical techniques. In the table, for each technique indicate:

- the chemical species being analysed;

- any chemical reagents other than water and the sample needed to carry out the analysis;
- whether or not a calibration curve is needed;
- sensitivity: low—concentrated (%) samples needed; high—very dilute (ppm) samples.

Technique	Species analysed	Other reagents	Calibration curve	Sensitivity
Colorimetry				
Volumetric analysis				
Atomic emission spectroscopy				

Q60.

Consider the following reactions.

I $A(g) + B(g) \rightarrow 2C(g)$; $\Delta H = +180$ kJ mol^{-1}

II $D(g) + 3E_2(g) \rightarrow 2F(g)$; $\Delta H = -90$ kJ mol^{-1}

III $2G(g) \rightarrow H(g) + I(g)$; $\Delta H = -180$ kJ mol^{-1}

From a comparison of the enthalpy change, ΔH, it can be deduced that:

A The activation energy of equation i > activation energy equation ii.

B The activation energy of equation i < activation energy equation ii.

C The activation energy of equation i = activation energy equation iii.

D No information about activation energy can be deduced from ΔH.

Q61.

In an endothermic reaction:

A the sum of the energy released when the bonds in the reactants break is greater than the sum of the energy absorbed when the bonds in the products form.

B the magnitude of the activation energy is always greater than the magnitude of ΔH.

C the products have a lower enthalpy than the reactants.

D energy from the reaction is released into the surroundings.

Q62.

Which of the following does not change the rate of collisions between particles in a reaction?

A addition of a catalyst

B increase in surface area

C increase in temperature

D decrease in the reaction volume

Q63.

Which of the following statements about a catalyst is true?

A Only a small amount of a catalyst is consumed in a reaction.
B A catalyst can occur as a reactant or product in the overall equation.
C A catalyst decreases the ΔH of the reaction.
D The proportion of molecules with sufficient energy to react is increased by a catalyst.

Q64.

A camper was having difficulty starting a wood fire. He took some pieces of wood and cut them into thinner pieces then blew on these as the first flames flickered. Which of the following techniques was he employing to increase the rate of reaction?

I increasing surface area
II increasing temperature
III increasing concentration of a reactant
IV addition of a catalyst

A I only.
B I and III.
C II and III.
D II and IV.

Q65.

Lumps of limestone, calcium carbonate, react readily with dilute hydrochloric acid. Four large lumps of limestone, mass 10.0 g, were reacted with 100 mL 0.1 M acid.

a Write a balanced equation to describe the reaction.
b Which reactant is in excess? Use a calculation to support your answer.
c Describe a technique that you could use in a school laboratory to measure the rate of the reaction.
d 10.0 g of small lumps of limestone will react at a different rate to four large lumps. Will the rate of reaction with the smaller lumps be faster or slower? Explain your answer in terms of collision theory.
e List two other ways in which the rate of this reaction can be altered. Explain your answer in terms of collision theory.

Q66.

The reaction A $\rightarrow$ B + C is found to have an activation energy of 630 kJ mol^{-1} and an enthalpy change of $\Delta H = -160$ kJ mol^{-1}.

a Draw an energy level diagram to describe the reaction.
b What would be the activation energy and the ΔH of the reverse reaction:
B + C $\rightarrow$ A.

Q67.

Consider the following equation: A + 3B $\rightleftharpoons$ 2C + 4D. The correct expression for the equilibrium constant is:

A $\dfrac{[A][3B]^3}{[2C]^2[4D]^4}$

B $\dfrac{[C]^2[D]^4}{[A][B]^3}$

C $\dfrac{[A][B]^3}{[C]^2[D]^4}$

D $\dfrac{[2C][4D]}{[A][3B]}$

Q68.

Consider the following equation: A + 3B $\rightleftharpoons$ 2C + 4D. Assuming all reactants and products are quoted in M units, the correct unit for the equilibrium expression is:

A no units

B M

C M^{-2}

D M^2

Q69.

A mixture of hydrogen gas and purple iodine vapour is sealed in a glass tube where it undergoes a reaction to form colourless hydrogen iodide gas. Which of the following could be used as a sign that equilibrium has been achieved?

A constant pressure

B constant colour

C constant mass

D constant volume

Q70.

Two samples of pure NO_2 are placed into separate sealed 1 L containers and allowed to reach equilibrium:

$$2NO_2 \rightleftharpoons N_2O_4$$

Mixture A consists initially of 2 mol pure NO_2. Mixture B consists initially of 1 mol NO_2. Both are at the same temperature. Consider the following values in each mixture at equilibrium:

I Ratio $[NO_2]/[N_2O_4]$ at equilibrium.

II Ratio $[NO_2]^2/[N_2O_4]$ at equilibrium.
III Ratio $[N_2O_4]/[NO_2]^2$ at equilibrium.
IV Time taken for equilibrium to be established.

Which will be different in the two mixtures?

A I only.
B I and II.
C II and III.
D I and IV.

Q71.

The solubility of precipitates can be increased by using competing equilibria. Silver ions readily form a white precipitate of silver chloride but form a soluble complex with ammonia:

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq) \qquad K = 1.6 \times 10^{-10}$$
$$AgCl(s) + 2NH_3(aq) \rightleftharpoons Ag(NH_3)_2^+(aq) + Cl^-(aq) \qquad K = 1.6 \times 10^{7}$$

In order to increase the solubility of the silver chloride precipitate, AgCl(s), it would be best to:

A add sodium chloride solution to the mixture.
B add ammonia solution to the mixture.
C add silver ions in the form of silver nitrate solution to the mixture.
D add H^+ ions to the mixture.

Q72.

Carbon monoxide and iodine pentoxide react to form iodine and carbon dioxide in the equilibrium reaction:

$$5CO(g) + I_2O_5(s) \rightleftharpoons I_2(g) + 5CO_2(g); \Delta H = -1175 \text{ kJ}$$

Use your knowledge of Le Chatelier's principle to predict the effect of the change (column 1) on the designated quantity (column 2). Complete the effect on the equilibrium (column 3) by using the terms decrease, increase, no change. Assume that the change listed is the only one taking place, e.g. if I_2 is added, the volume and the temperature are kept constant.

Change	Quantity	Effect
Increase *T*	*K*	
Decrease *T*	amount $I_2O_5(s)$	
Add $I_2(g)$	*K*	
Add $CO_2(g)$	amount $I_2(g)$	
Double volume	concentration CO_2	

Remove CO	amount CO_2
Add catalyst	$I_2(g)$
Add inert gas Ar	K

Q73.

The equilibrium for the reaction of nitrogen dioxide to form nitrogen tetroxide is shown:

$$(g) \rightleftharpoons N_2O_4(g);\ K = 170 \text{ at } 298\ K$$

A mixture of the gases was found to contain 0.015 M NO_2 and 0.025 M N_2O_4.

a What are the correct units for the equilibrium constant, *K*, for the equation?

b Calculate the reaction quotient for the mixture shown.

c Is the mixture at equilibrium? If no, should more N_2O_4 or NO_2 be formed for the mixture to reach equilibrium?

Q74.

If hydrogen iodide is heated it will decompose to form iodine and hydrogen. In a sealed container at 425°C the following equilibrium was obtained:

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g);\ K = 55.64 \text{ at } 425°C$$

Initially 2.00 mol H_2 and 2.00 mol I_2 are placed in a 1 L flask with no hydrogen iodide is present.

a If *x* moles of hydrogen react to reach equilibrium, how many moles of iodine are consumed and how many moles of hydrogen iodide are formed?

b What is the equilibrium concentration of hydrogen, iodine and hydrogen iodide, in terms of *x*?

c Substitute the values of the equilibrium concentrations of hydrogen, iodine and hydrogen iodide into the equilibrium expression and determine the number of moles of hydrogen iodide at equilibrium.

Q75.

The solubility of carbon dioxide in natural systems is affected by a number of competing equilibria

Equation 1 $CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

Equation 2 $H_2CO_3(aq) + Ca(OH)_2(aq) \rightleftharpoons CaCO_3(s) + 2H_2O(l)$

Equation 3 $CaCO_3(s) + CO_2(g) + H_2O(l) \rightleftharpoons Ca^{2+}(aq) + 2HCO_3^-(aq)$

a The concentration of carbon dioxide in the air has increased over the last 200 years. What is the source of this increased carbon dioxide?

b Use the equation(s) above and your knowledge of Le Chatelier's principle to explain how the oceans can act as a carbon dioxide sink

c Use the equation(s) above to explain the dissolution of limestone in carbon dioxide rich waters in caves.

Q76.

Blood plasma has a hydrogen ion concentration of 3.0×10^{-7} M. The pH of the blood plasma is closest to:

A 3.7
B 6.5
C 7
D 7.5

Q77.

A sample of lactic acid, a weak monobasic acid, was found to have a pH = 2.43 at 25°C. The concentration of OH^- of the acid was:

A 3.71×10^{-3} M
B 3.71×10^{-11} M
C 2.70×10^{-12} M
D 4.12×10^{-15} M

Q78.

The pH of 0.1 M HCl(aq) is approximately 1.0. The pH of 0.1 M of the diprotic acid malonic acid, $CH_2(COOH)_2$, is 1.9. The pH of the malonic acid solution is higher because:

A HCl ionises to form more H_3O^+ ions in water than malonic acid.
B malonic acid has more hydrogen ions than HCl.
C HCl is a weaker acid than malonic acid.
D malonic acid is amphiprotic in solution.

Q79.

The pH values of a series of 0.1 M solutions of a series of chloroxacids, HClO, $HClO_2$, $HClO_3$ and $HClO_4$, were determined. pH was found to decrease in the following order: HClO > $HClO_2$ > $HClO_3$ > $HClO_4$. From these data it can be inferred that:

A HClO is a stronger acid than $HClO_4$.
B the K_a of $HClO_2$ is less than that of $HClO_3$.
C the lower the oxidation number of Cl the stronger the acid.
D the greater the number of oxygen atoms in the acid, the stronger the O-H bond.

Q80.

A sample of 0.10 M acetic acid was found to have a pH of 2.89. The percentage dissociation of the 0.10 M acetic acid sample was:

A 0.21%
B 0.59%
C 1.3%
D 3.5%

Q81.

The pH of 1.0 M methanoic acid (a monobasic acid, HCOOH) is 1.25. The pH of 1.0 M HNO_3 is 0.

a What is the concentration of H_3O^+ in a solution of methanoic acid, pH = 1.25?

b Why is the pH of the methanoic acid higher than that of nitric acid? Include relevant equations in your answers

Q82.

a Write ionic equations to explain why:
  i ammonium chloride has a pH less than 7.
  ii sodium sulfide has a pH greater than 7.

b Sodium carbonate is readily soluble in water while calcium carbonate is insoluble. A solution of sodium carbonate was found to have a pH of 8. A quantity of calcium nitrate was added to the solution. Using ionic equations, explain the effect the addition of calcium nitrate would have on the pH of the solution.

Q83.

Propoanoic acid (abbreviated to HPr) is a weak monobasic acid used to retard the growth of mould in foods. A 0.100 M solution was prepared. The acidity constant for the acid is:

$$K_a = \frac{[H_3O^+][Pr^-]}{[HPr]} = 1.30 \times 10^{-5}\ \text{M}$$

a Assuming *x* moles of HPr dissociates, complete the following table.

	[HPr]	$[H_3O^+]$	$[Pr^-]$
Initial concentration (M)	0.100	0	0
Change in concentration (M)	$-x$		

Equilibrium concentration (M)	$0.100 - x \approx 0.100$

b Determine $[H_3O^+]$.

c Determine the pH of the solution.

d Determine the percentage dissociation of HPr.