

PRACTICE WORKSHEET + FULL ANSWER KEY

Moles, reactions and solutions

Senior secondary chemistry. Work through the sections in order, or choose the one that matches this week's topic.

QUESTIONS	SECTIONS	ANSWER KEY	WORKING SPACE
104	10	At the back	Ruled

Name _____ Class _____ Date _____

Before you start

A balanced equation gives particle and mole ratios, not a direct mass ratio. Convert mass to moles first, use the stoichiometric ratio, then convert to the quantity requested.

WORKED EXAMPLE

Find the amount in 9 g of water. Use molar mass 18 g/mol.

- 1 Use $n = m/M$.
- 2 $n = 9/18 = 0.5$.
- 3 There are 0.5 mol of water.

Use this method as your model. Set out your own working the same way.

Common mistakes to avoid

- Changing subscripts to balance an equation
- Using millilitres in a concentration formula that requires litres

There are 104 questions across 10 sections. Write in the ruled space, then mark your work against the answer key at the back. Topics covered: Balanced equations, Amount of substance, Concentration.

SECTION A Atoms, bonding and the periodic table

Start from the electron arrangement, then explain the property it produces.

1

State the relative charge and mass of a proton, a neutron and an electron.

3

Define an isotope.

2

An atom has 11 protons and 12 neutrons. Give its mass number and name the element.

4

Write the electronic configuration of a chlorine atom, $Z = 17$.

5

Explain why group 1 metals become more reactive down the group.

6

State and explain the bonding in sodium chloride.

7

State the bonding in methane.

8

Explain why ionic compounds conduct when molten but not when solid.

9

Explain why diamond is hard but graphite is soft.

10

Write the formula of magnesium chloride.

11

Write the formula of calcium carbonate.

SECTION B Equations and reaction types

Balance by changing coefficients only. Never change a subscript.

12

Balance: $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$.

13

Balance: $\text{Na} + \text{Cl}_2 \rightarrow \text{NaCl}$.

14

Balance: $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$.

15

Balance: $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$.

16

Balance: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$.

17

Write a balanced equation for hydrochloric acid neutralised by sodium hydroxide.

18

Name the salt formed when sulfuric acid reacts with potassium hydroxide.

19

Explain why atoms are conserved in a chemical reaction.

20

State what is observed when magnesium reacts with dilute hydrochloric acid.

21

Explain why changing a subscript is not a valid way to balance an equation.

22

Classify as oxidation or reduction: $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$.

SECTION C Amount of substance

Convert mass to moles first, apply the ratio from the balanced equation, then convert back.

- 23 Find the amount in 9 g of water, using $M = 18$ g/mol.

- 24 Find the amount in 22 g of CO_2 , using $M = 44$ g/mol.

- 25 Find the mass of 0.2 mol of NaCl , using $M = 58.5$ g/mol.

- 26 Find the mass of 3 mol of water.

- 27 Calculate the relative formula mass of CaCO_3 , using Ca 40, C 12 and O 16.

- 28 Calculate the relative formula mass of H_2SO_4 , using H 1, S 32 and O 16.

- 29 For $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, find the moles of NH_3 from 6 mol of H_2 with excess N_2 .

- 30 For $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, find the mass of water from 4 g of hydrogen with excess oxygen.

- 31 Find the number of particles in 0.5 mol, using 6.02×10^{23} per mole.

- 32 Calculate the percentage by mass of oxygen in CaCO_3 .

- 33 A reaction has a theoretical yield of 8.0 g and an actual yield of 6.4 g. Find the percentage yield.

SECTION D Concentration, gases and acids

Volumes must be in litres for $n = cV$. Check the unit before you substitute.

- 34 Find the concentration of 0.3 mol dissolved in 1.5 L.

- 35 Find the amount of solute in 250 mL of 0.4 mol/L solution.

- 36 Find the volume of 0.5 mol/L solution containing 0.05 mol.

- 37 Convert 250 cm^3 to litres.

- 38 25.0 cm^3 of NaOH needs 20.0 cm^3 of 0.100 mol/L HCl . Find the NaOH concentration.

39

Find the volume of 2 mol of gas at room conditions, using 24 L/mol.

40

Find the amount of gas in 6 L at room conditions.

41

Explain why a concentration formula requires litres rather than millilitres.

42

Explain the difference between a concentrated acid and a strong acid.

43

State the pH range of an acid, a neutral solution and an alkali.

44

Explain what happens to the pH of a strong acid diluted tenfold.

SECTION E Practical chemistry and rates

Describe the apparatus, the safety control and the reason the result can be trusted.

45

Name the apparatus used to measure 25.0 cm³ accurately in a titration.

46

State one safety precaution when heating a solution in a test tube.

47

Explain why a titration is repeated until concordant results are obtained.

48

A student records 24.6, 24.8 and 27.1 cm³. Which values should be averaged?

49

Name a test for hydrogen gas and give the result.

50

Name a test for carbon dioxide and give the result.

51

Explain why a reaction speeds up at a higher temperature.

52

Explain the effect of a catalyst on a reaction.

53

Explain why a powdered solid reacts faster than a single lump.

54

A student calculates a percentage yield above 100%. Suggest one reason.

SECTION F Rates and equilibrium

Explain rate changes through collisions; explain equilibrium shifts through opposing the change.

55 State four factors that affect the rate of a reaction.

56 Define activation energy.

57 Explain the effect of increasing concentration on rate.

58 Describe the shape of a graph of gas volume against time.

59 State how to find the rate from such a graph.

60 State Le Chatelier's principle.

61 For the exothermic reaction $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$, state the effect of raising the temperature.

62 For the same reaction, state the effect of increasing the pressure.

63 State the effect of a catalyst on the position of equilibrium.

64 Define a dynamic equilibrium.

SECTION G Organic chemistry

Count the carbons, then apply the general formula for the family.

65 State the general formula of the alkanes.

66 Write the formula of ethane.

67 Write the formula of propene.

68 State the general formula of the alkenes.

69 Name the test that distinguishes an alkene from an alkane and give the result.

70 Name the products of complete combustion of a hydrocarbon.

71

Name a product of incomplete combustion and state one hazard.

72

Define cracking and give one reason it is used.

73

Name the functional group in ethanol.

74

Explain what a homologous series is.

SECTION H Redox and electrochemistry

Track the electrons: loss is oxidation, gain is reduction.

75

Define oxidation in terms of electrons.

76

Define reduction in terms of electrons.

77

In $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, state what is oxidised.

78

Name the electrode at which oxidation occurs during electrolysis.

79

Name the products of the electrolysis of molten lead bromide.

80

Explain why an ionic compound must be molten or dissolved to be electrolysed.

81

Name the product at the cathode when electrolysis aqueous sodium chloride.

82

Order these most reactive first: copper, magnesium, zinc.

83

Explain why magnesium displaces copper from copper sulfate solution.

84

Write the half equation for chlorine forming at an anode.

SECTION I Energetics

Bond breaking takes energy in; bond making gives energy out.

85 Define an exothermic reaction.

86 Define an endothermic reaction.

87 Give one example of an exothermic process.

88 Give one example of an endothermic process.

89 On an energy profile for an exothermic reaction, where do the products sit?

90 State what the activation energy looks like on an energy profile.

91 Explain the effect of a catalyst on an energy profile.

92 Explain why bond breaking is endothermic.

93 A reaction breaks bonds needing 1,200 kJ and forms bonds releasing 1,500 kJ. Find the overall change.

94 State the units usually used for an enthalpy change.

SECTION J Mixed calculations and practical review

Write the formula, substitute with units, then evaluate.

95 Find the amount in 5.85 g of NaCl, using $M = 58.5$ g/mol.

96 Find the mass of 0.25 mol of CaCO_3 , using $M = 100$ g/mol.

97 Find the concentration when 4 g of NaOH ($M = 40$ g/mol) makes 500 mL of solution.

98 Find the volume of 0.5 mol of gas at room conditions, using 24 L/mol.

99 Find the empirical formula of a compound that is 40% C, 6.7% H and 53.3% O by mass.

100 Name the apparatus used to separate two liquids with different boiling points.

101

Explain why chromatography separates a mixture of dyes.

102

Calculate the R_f value when a spot moves 4.5 cm and the solvent front moves 9.0 cm.

103

State one safety precaution when using concentrated acid.

104

Explain why a reaction mixture is often stirred.

Answer key

Work each section before you check it. These are model answers: an equivalent method, or a different well-supported interpretation, can be just as correct.

SECTION A - ATOMS, BONDING AND THE PERIODIC TABLE

1. Proton: charge +1, mass 1. Neutron: charge 0, mass 1. Electron: charge -1, mass about 1/1836.
2. Mass number 23; the element is sodium.
3. Atoms of the same element with the same number of protons but different numbers of neutrons.
4. 2, 8, 7.
5. The outer electron is further from the nucleus and better shielded, so it is lost more easily.
6. Ionic. Sodium transfers an electron to chlorine and the oppositely charged ions attract in a lattice.
7. Covalent; the atoms share pairs of electrons.
8. Ions can move when molten; in the solid they are held in fixed positions in the lattice.
9. Diamond has a rigid covalent network; graphite has layers held together by weak forces that can slide.
10. MgCl_2 .
11. CaCO_3 .

SECTION B - EQUATIONS AND REACTION TYPES

12. $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$.
13. $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$.
14. $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$.
15. $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$.
16. It is already balanced as written.
17. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$.
18. Potassium sulfate.
19. Atoms are rearranged into new combinations; ordinary chemical reactions do not create or destroy them.
20. Effervescence as hydrogen is released, the magnesium dissolves and the mixture warms.
21. It changes the identity of the substance; only the coefficients may be adjusted.
22. Oxidation, because electrons are lost.

SECTION C - AMOUNT OF SUBSTANCE

23. 0.5 mol.
24. 0.5 mol.
25. 11.7 g.
26. 54 g.
27. 100.
28. 98.
29. 4 mol of NH_3 .
30. 36 g, because 2 mol of H_2 produces 2 mol of water.

31. About 3.01×10^{23} particles.

32. 48%, from 48/100.

33. 80%.

SECTION D - CONCENTRATION, GASES AND ACIDS

34. 0.2 mol/L.
35. 0.1 mol.
36. 0.1 L, which is 100 mL.
37. 0.25 L.
38. 0.0800 mol/L, because 0.00200 mol reacts 1:1 in 0.0250 L.
39. 48 L.
40. 0.25 mol.
41. The unit mol/L needs the volume in litres; using millilitres makes the answer 1,000 times too large.
42. Concentrated describes how much acid there is per litre; strong describes how fully it ionises in water.
43. Below 7, exactly 7, and above 7.
44. The pH rises by about 1 unit, because the hydrogen ion concentration falls by a factor of ten.

SECTION E - PRACTICAL CHEMISTRY AND RATES

45. A pipette, used with a pipette filler.
46. Wear eye protection and heat gently with the tube pointing away from people.
47. Concordant titres show the technique is consistent, so the mean is reliable.
48. 24.6 and 24.8, giving 24.7 cm^3 ; 27.1 is anomalous and excluded with a reason recorded.
49. A lit splint gives a squeaky pop.
50. Limewater turns milky, or cloudy.
51. Particles move faster and more collisions exceed the activation energy, so successful collisions become more frequent.
52. It provides an alternative pathway with a lower activation energy and is not used up.
53. A larger surface area gives more collisions per second between reactant particles.
54. The product was not fully dried, so the mass included water or another impurity.

SECTION F - RATES AND EQUILIBRIUM

55. Temperature, concentration (or pressure for gases), surface area and a catalyst.
56. The minimum energy colliding particles must have for a reaction to occur.

57. More particles per unit volume means more frequent collisions, so the rate rises.
58. It rises steeply at first, then levels off as a reactant is used up.
59. Take the gradient of the curve; the initial rate is the gradient at time zero.
60. If a change is applied to a system at equilibrium, the position of equilibrium shifts to oppose that change.
61. Equilibrium shifts to the left, reducing the yield of ammonia.
62. Equilibrium shifts right, towards the fewer moles of gas, increasing the yield.
63. None. It only speeds up the approach to equilibrium.
64. The forward and reverse reactions continue at equal rates, so the concentrations stay constant.

SECTION G - ORGANIC CHEMISTRY

65. C_nH_{2n+2} .
66. C_2H_6 .
67. C_3H_6 .
68. C_nH_{2n} .
69. Bromine water; it is decolourised by an alkene.
70. Carbon dioxide and water.
71. Carbon monoxide, which is toxic because it reduces oxygen transport in the blood.
72. Breaking long hydrocarbons into shorter ones; it produces more useful short-chain fuels and alkenes.
73. The hydroxyl group, -OH.
74. A family of compounds with the same general formula and functional group, differing by CH_2 , with gradually changing properties.

SECTION H - REDOX AND ELECTROCHEMISTRY

75. Loss of electrons.
76. Gain of electrons.
77. Zinc, because it loses two electrons.
78. The anode, the positive electrode.
79. Lead at the cathode and bromine at the anode.
80. The ions must be free to move so they can carry the current.
81. Hydrogen, because sodium is more reactive than hydrogen.
82. Magnesium, zinc, copper.
83. Magnesium is more reactive, so it loses electrons more readily and takes copper's place.
84. $2Cl^- \rightarrow Cl_2 + 2e^-$.

SECTION I - ENERGETICS

85. One that transfers energy to the surroundings, so the temperature rises.

86. One that takes in energy from the surroundings, so the temperature falls.
87. Combustion, neutralisation or respiration.
88. Thermal decomposition, or dissolving ammonium nitrate.
89. Below the reactants.
90. The height from the reactant level up to the peak.
91. It lowers the peak, giving a lower activation energy, without changing the reactant or product levels.
92. Energy must be supplied to overcome the attraction holding the atoms together.
93. -300 kJ, so the reaction is exothermic.
94. Kilojoules per mole, kJ/mol.

SECTION J - MIXED CALCULATIONS AND PRACTICAL REVIEW

95. 0.1 mol.
96. 25 g.
97. 0.2 mol/L.
98. 12 L.
99. CH_2O .
100. A fractional distillation set-up with a fractionating column.
101. Each dye has a different attraction to the paper and the solvent, so they travel different distances.
102. 0.5.
103. Wear eye protection and gloves, and add acid to water rather than water to acid.
104. It keeps the concentration even so reactant particles meet throughout the mixture.

Review and next steps

Mark your work against the answer key, record your score for each section, then answer the three questions below. Bring this page to your next lesson so your tutor can start from what actually blocked you.

Which question did you find hardest, and what exactly stopped you?

Rewrite the corrected working for that question here.

Which section do you want to practise again before your next lesson?

Score record

SECTION	QUESTIONS	NUMBER CORRECT	PRACTISE AGAIN?
A. Atoms, bonding and the periodic table	11		
B. Equations and reaction types	11		
C. Amount of substance	11		
D. Concentration, gases and acids	11		
E. Practical chemistry and rates	10		
F. Rates and equilibrium	10		
G. Organic chemistry	10		
H. Redox and electrochemistry	10		
I. Energetics	10		
J. Mixed calculations and practical review	10		
Total	104		

Keep going with one-to-one support.

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