UNIT 3

AMOUNT OF SUBSTANCE AND MEASUREMENT

Answers

Lesson 1 – Why are practicals important?

P Act	tivity 1.1: Understand risks and safety precautions in the laboratory
precaution	nould simply be encouraged to take time to discuss and present any of the laboratory safety s from the list above. It may be necessary to distribute colouring pencils and /or pens in order to tudents to make an effort with their poster. The best posters should be displayed in the laboratory.
P	Summary Activity 1.2: units of temperature
- 298	lvin, degrees celsius (and degrees Farenheit) 8 K, 373 K, 0 K °C, 327 °C, -173 °C
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Test your knowledge 1.3: Interconverting important units in Chemistry

(a) (i) 25000 g; (ii) 3200 g; (iii) 340 g

(b) (i) 2.5 x 10⁻⁵ m³; (ii) 3.2 x 10⁻³ m³; (iii) 3.4 x 10⁻⁴ m³, (d) 1.5 x 10⁻⁴ m³, (e) 0.12 m³

- (c) (i) 250 dm³; (ii) 3200 dm³; (iii) 0.025 dm³; (iv) 0.15 dm³; (v) 6.2 x 10⁻³ dm³
- (d) (i) 2.5 x 10⁵ cm³, (b) 3.2 x 10⁶ cm³, (c) 400 cm³, (d) 15 cm³, (e) 6200 cm³

Lesson 2 – What is a base quantity and what is a derived quantity?

Test your knowledge 2.1: Using base and derived quantities

(a) Force (= mass x acceleration) kgms⁻²

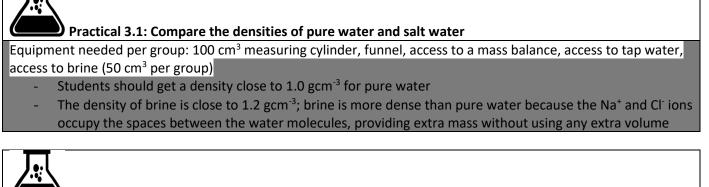
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- (b) Work done (= pressure x volume) kgm²s⁻²
- (c) Power (= voltage x current) kgm²s⁻³
- (d) Momentum (= mass x velocity) kgms⁻¹
- (e) Rate of reaction (= concentration / time) molm $^{-3}$ s⁻¹

Test your knov	vledge 2.2: Measuring Volum	nes
Instrument	Advantage	Disadvantage
Pipette	very accurate	can only measure one volume
Volumetric flask	very accurate	can only measure one volume
Burette	Can measure any volume up to 50 cm ³	Cannot measure the total volume present, it can only deliver a volume
Measuring cylinder	Easy to use	Not very accurate

UNIT 3 – AMOUNT OF SUBSTANCE AND MEASUREMENT

Lesson 3 – What is density and how can we measure it?



Practical 3.2: Measure the density of sand

Equipment needed per group: 2 x 100 cm³ measuring cylinders, access to a mass balance, access to tap water, access to sand (around 20 g per group), access to a spoon

- students should get a density close to 1.5 gcm⁻³
- sand must be denser than water because it does not float on water
- The error in the measurement of volume is the biggest error, as measuring cylinders are not very accurate

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Test your knowledge 3.3: Using Avogadro's number

(a) 0.0042 or 4.2 x 10⁻³ (b) 1.5 x 10²³ (c) 0.05 (d) 1.2 x 10²² (e) 15

Lesson 4 – How can we work out how many moles we have in a sample?

Test your knowledge 4.1: Deducing relative molecular masses and relative formula masses (a) (i) 12.0; (ii) 16.0; (iii) 35.5; (iv) 23.0; (v) 1.0; (vi) 24.3 (b) (i) 32.0; (ii) 44.0; (iii) 71.0; (iv) 36.5; (v) 16.0; (vi) 18.0 (c) (i) 58.5; (iii) 106.0; (iv) 40.3; (v) 95.3; (v) 58.3

a) (i) 0.1; (ii) 0.078; (iii) 5450; (iv) 0.16, (v) 0.022 b) (ii) 3.55 g; (ii) 14.9 g; (iii) 5.56 g; (iv) 39900 g or 39.9 kg; (v) 6.85 g c) (i) 28 g/mol; (ii) 40 g/mol; (iii) 160 g/mol; (iv) 28 g/mol; (v) 249.6 g/mol

Lesson 5 - How can we work out how many moles we have in a solution?

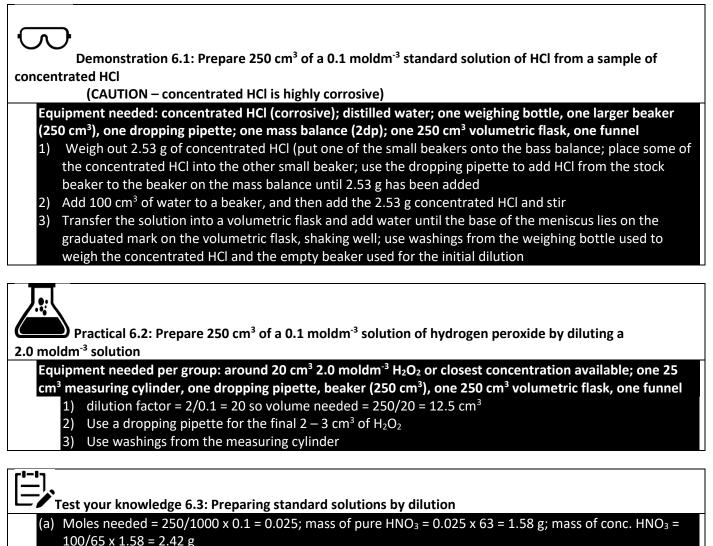
Test your knowledge 5.1: Using moles, molarity and aqueous volume	
 a) (i) 0.0025; (ii) 0.008; (iii) 0.015; (iv) 0.0025; (v) 0.0052 b) (i) 2.5 moldm⁻³; (ii) 0.4 moldm⁻³; (iii) 0.12 moldm⁻³; (iv) 0.1 moldm⁻³; (v) 2 moldm⁻³ c) (i) 6.0 moldm⁻³; (ii) 0.63 moldm⁻³; (iii) 2.7 moldm⁻³ (iv) 0.8 moldm⁻³, (v) 2.9 moldm⁻³ 	

Test your knowledge 5.2: Preparing Standard Solutions

(a) moles needed = $250/1000 \times 0.1 = 0.025$; m_r = 106 so mass needed = $0.025 \times 106 = 2.65$ g (b) moles needed = $250/1000 \times 0.1 = 0.025$; m_r = 174 so mass needed = $0.025 \times 174 = 4.35$ g

Practical 5.3: Prepare 250 cm ³ of 0.1 moldm ⁻³ standard solutions of sodium chloride (NaCl) and sugar (C ₁₂ H ₂₂ O ₁₁)
Equipment needed per group: 250 cm ³ beaker, distilled water bottle, spatula, stirring rod, funnel, 250 cm ³ volumetric flask, weighing boat, access to 2 dp mass balance, access to NaCl, access fo sugar
- Mass of salt needed = $58.5 \times 0.25 \times 0.1 = 1.46 \text{ g}$
 Mass of sugar needed = 342 x 0.25 x 0.1 = 8.55 g
It ma be advisable to prepare the standard solution of NaCl together, as a class, with the teacher leading from the
front showing the key steps, before allowing the students to prepare the sugar solution independently

Lesson 6 – How can we prepare standard solutions by diluting concentrated solutions?



(b) Moles of NaOH = $5/1000 \times 6 = 0.03$; total volume of diluted solution = $0.03/0.1 = 0.3 \text{ dm}^3 = 300 \text{ cm}^3$; So $300 - 5 = 295 \text{ cm}^3$ of water must be added

UNIT 3 – AMOUNT OF SUBSTANCE AND MEASUREMENT

Lesson 7 – How can calculate the moles present in a gaseous sample?



Summary Activity 7.1: The Gas Laws

- Because the particles are far apart and there are no forces between the particles

- The typical pressure exerted on the earth'y surface by its atmosphere; 100 kPa (also known as 1 atm)

- $P_1V_1/T_1 = P_2V_2/T_2$
- Boyle's Law, Charles' Law and Gay-Lussac's Law (any two of these can be used to derive the combined gas law)



Test your knowledge 7.2: Using Avogadro's Law

(a) 2.4 dm³, (b) 7.2 dm³, (c) 24 dm³, (d) 0.5, (e) 0.005

Note: it doesn't matter what the gas is; the gas laws apply equally to all gases



Demonstration 7.3: Measure the volume of a gas

Equipment needed: conical flask, bung which fits conical flask and has delivery tube attached, gas syringe connectible to delivery tube (or trough of water and 100 cm³ measuring cylinder), 50 cm³ measuring cylinder, access to 2.0 moldm⁻³ HCl and marble chips

Using the measuring cylinder, pour around 50 cm³ of 2.0 moldm⁻³ HCl into the conical flask; ensure that the delivery tube with the bung is connected to the syringe; add 0.25 g – 0.30 g of marble chips and quickly replace the bung; the plunger in the syringe will move and the volume of gas can be measured (expect 50 – 70 cm³ of gas) Record the atmospheric temperature and inform the class

Moles of gas = n = PV/RT; Pressure = 100,000 Pa, R = 8.31; T = (eg) 20 °C = 293 K (use class measurement); V = (eg) 65 cm³ = 6.5 x 10⁻⁵ m³ (use class measurement); number of moles of gas produced = (100,000 x 6.5 x 10⁻⁵)/ (8.31 x 293) = 2.7 x 10⁻³ moles (this is an example using V = 65 cm³) and T = 20 °C)

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Test your knowledge 7.4: Using the ideal gas equation

(a) (i) 1.9 mol; (ii) 0.048 mol; (iii) 0.0022 mol; (iv) 4.0 mol; (v) 0.0024 mol

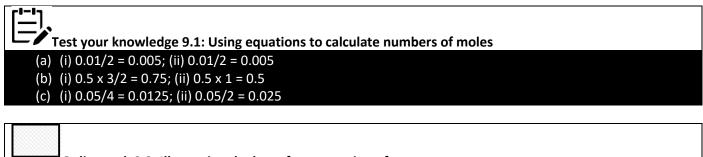
- (b) (i) 1.2 dm³, (ii) 6.2 dm³, (iii) 9.1 dm³, (iv) 2.5 dm³, (v) 11 dm³
- (c) (i) 62 g, (ii) 2.1 g, (iii) 0.62 g, (iv) 280 g, (v) 0.11 g

Lesson 8 – What is an empirical formula and how is it different from a molecular formula or a unit formula?

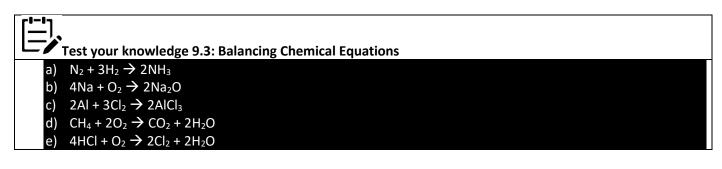
 Summary Activity 8.1: Unit formula and molecular formula
 Molecular formula: number of atoms of each element in one molecule: eg C₆H₁₂O₆ or CO₂ Unit formula: simplest ratio of each particle in the compound: eg NaCl, Ca(OH)₂

_	Test Your Knowledge 8.2: Empirical Formulae
a)	ef: $62.08/12:10.34/1:27.58/16 = 5.17:10.34:1.72 = 3:6:1$ so ef = C ₃ H ₆ O; efm = 58 and rmm = 58 so n = 58/58
b)	ef: $22.02/12:4.59/1:27.73.39/79.9 = 1.84:4.59:0.92 = 2:5:1$ so ef = C_2H_5Br
c)	ef: 84.21/12: 15.79/1 = 7.01:15.79 = 1:2.25 = 4:9 so ef = C_4H_9 ; efm = 57 and rmm = 114 so n = 114/57 = 2 so mf
	$= C_8 H_{18} O$
d)	$7.8 - 0.6 = 7.2$ g of C; ef: $72/12$: $6/1 = 6$: $6 = 1$:1 so ef = CH; efm = 13 and rmm = 78 so n = $78/13 = 6$ so mf = C_6H_6
e)	ef: 3.36/55.8:1.44/16 = 0.06:0.09 = 1:1.5 = 2:3 so ef = Fe ₂ O ₃
f)	ef: 48.4/16:24.3/32.1:21.2/14:6/1/1 = 3.03:0.76:1.51:6.1 = 4:1:2:8 so ef = O ₄ SN ₂ H ₈ ; unit formula = (NH ₄) ₂ SO ₄

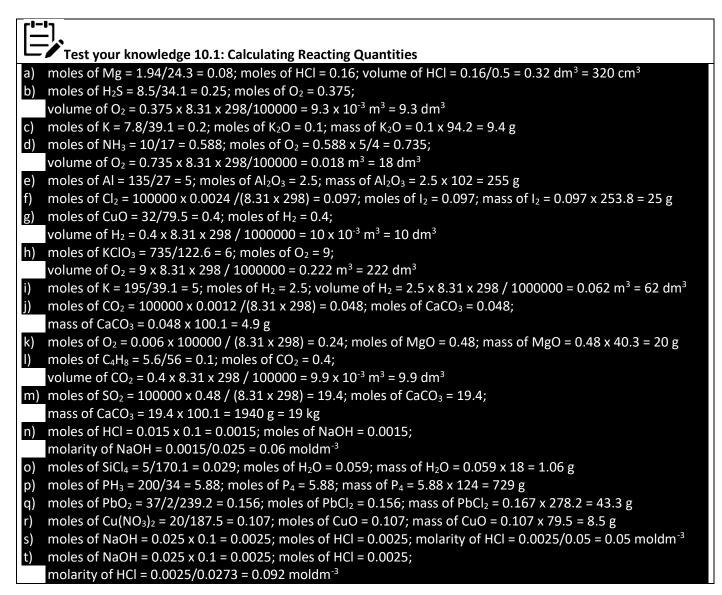
Lesson 9 – What are chemical equations and why are they useful?



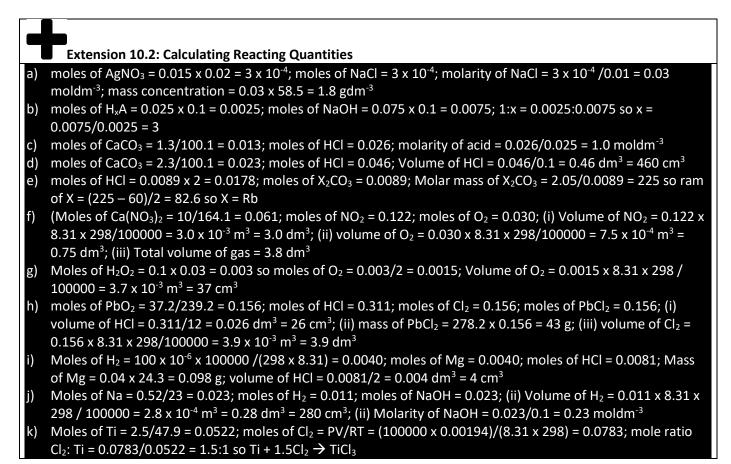
Online task 9.2: Illustrating the law of conservation of mass



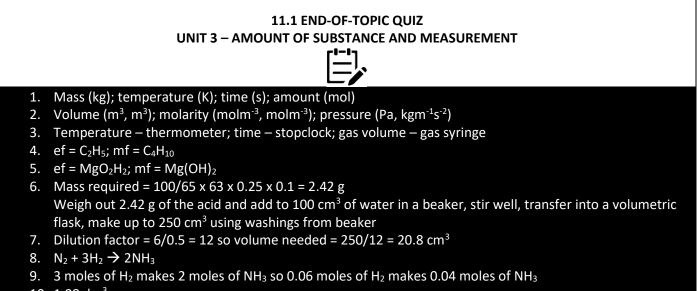
Lesson 10 - How can we use chemical equations to predict reacting quantities?



UNIT 3 – AMOUNT OF SUBSTANCE AND MEASUREMENT



Lesson 11 – What have I understood about Amount of Substance and Measurement?



- 10. 1.08 dm³
- 11. 53.3 cm³
- 12. 1.49 dm³