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Stoichiometry: Molar Masses, Formulas & Calculations

Printable Flashcards — Pre-Med Chemistry

Moles, Avogadro's number, atomic/molar masses of 50+ compounds, diatomic elements, polyatomic ions, formula writing, limiting reagents, percent yield, molarity, dilution, gas laws, and empirical vs molecular formulas.

193 cards — Print double-sided, flip on long edge, then cut along dashed lines.

193 cards — Printable Flashcards

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1

Stoichiometry in one line: what are we doing?

2

The 4-step stoichiometry routine that saves you every time:

3

Trap: you can balance an equation by changing subscripts in formulas. True or false?

4

What do coefficients in a balanced equation actually mean?

5

Key conversion: moles from grams =

6

Moles = mass / {{c1::molar mass}}.

7

Avogadro's number is basically...

8

Trap: 1 mole = 6.02×10^{23} grams. True or false?

2

1) Balance. 2) Convert to moles. 3) Use mole ratio. 4) Convert to wanted units.

1

Using a balanced equation to convert between amounts of reactants and products.

4

Mole ratios (and also molecule ratios).

3

False.

6

Moles = mass / molar mass.

5

moles = mass (g) / molar mass (g/mol).

8

False.

7

6.02×10^{23} particles per mole.

9

Particles to moles conversion:

10

Moles to particles conversion:

11

What is molar mass (in plain words)?

12

How do you find molar mass of a compound quickly?

13

Trap: coefficients change the molar mass. True or false?

14

Big pre-med trick: always write units at every step because...

15

Quick formula you should know cold: density =

16

If you have a solution and you know molarity (M), moles =

10

particles = moles x (6.02×10^{23}).

9

moles = particles / (6.02×10^{23}).

12

Add atomic masses for every atom in the formula (watch subscripts + parentheses).

11

The mass of 1 mole of a substance (in g/mol).

14

It tells you if your conversion makes sense.

13

False.

16

moles = M x volume (L).

15

mass / volume (d = m/V).

17

Molarity: M = {{c1::moles}}/{{c2::liters}}.

18

Trap: if volume is in mL, you can plug it directly into molarity formula. True or false?

19

Atomic mass of Hydrogen (H) is about (g/mol):

20

Atomic mass of Carbon (C) is about (g/mol):

21

Atomic mass of Nitrogen (N) is about (g/mol):

22

Atomic mass of Oxygen (O) is about (g/mol):

23

Atomic mass of Fluorine (F) is about (g/mol):

24

Atomic mass of Sodium (Na) is about (g/mol):

18

False.

17

Molarity: M = moles/liters.

20

12

19

1

22

16

21

14

24

23

23

19

25

Atomic mass of Magnesium (Mg) is about (g/mol):

26

Atomic mass of Aluminum (Al) is about (g/mol):

27

Atomic mass of Silicon (Si) is about (g/mol):

28

Atomic mass of Phosphorus (P) is about (g/mol):

29

Atomic mass of Sulfur (S) is about (g/mol):

30

Atomic mass of Chlorine (Cl) is about (g/mol):

31

Atomic mass of Potassium (K) is about (g/mol):

32

Atomic mass of Calcium (Ca) is about (g/mol):

26

27

25

24

28

31

27

28

30

35.5

29

32

32

40

31

39

33

Atomic mass of Iron (Fe) is about (g/mol):

34

Atomic mass of Copper (Cu) is about (g/mol):

35

Atomic mass of Zinc (Zn) is about (g/mol):

36

Atomic mass of Bromine (Br) is about (g/mol):

37

Atomic mass of Iodine (I) is about (g/mol):

38

Atomic mass of Silver (Ag) is about (g/mol):

39

Atomic mass of Barium (Ba) is about (g/mol):

40

The 4 numbers you use constantly in molar mass questions:

34

63.5

33

56

36

80

35

65

38

108

37

127

40

H=1, C=12, N=14, O=16.

39

137

41

Quick atomic mass set: H={{c1::1}}, C={{c2::12}}, N={{c3::14}}, O={{c4::16}}, Na={{c5::23}}, Cl={{c6::35.5}}.

42

Diatomic elements you must remember (when they're 'alone'):

43

Trap: oxygen gas is written as O. True or false?

44

Elemental nitrogen in equations is written as:

45

Elemental chlorine in equations is written as:

46

Halogens as elements are diatomic: {{c1::F2}}, {{c2::Cl2}}, {{c3::Br2}}, {{c4::I2}}.

47

Molar mass of H2O (g/mol):

48

Molar mass of H2O2 (g/mol):

42

H2, N2, O2, F2, Cl2, Br2, I2.

41

Quick atomic mass set: H=1, C=12, N=14, O=16, Na=23, Cl=35.5.

44

N2

43

False.

46

Halogens as elements are diatomic: F2, Cl2, Br2, I2.

45

Cl2

48

34

47

18

49

Molar mass of CO_2 (g/mol):

50

Molar mass of CO (g/mol):

51

Molar mass of O_2 (g/mol):

52

Molar mass of N_2 (g/mol):

53

Molar mass of NH_3 (g/mol):

54

Molar mass of CH_4 (g/mol):

55

Molar mass of C_2H_6 (g/mol):

56

Molar mass of C_3H_8 (g/mol):

50

28

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49

44

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52

28

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51

32

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54

16

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53

17

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56

44

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55

30

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57

Molar mass of C4H10 (g/mol):

58

Molar mass of C2H6O (ethanol) (g/mol):

59

Molar mass of C2H4O2 (acetic acid) (g/mol):

60

Molar mass of C6H12O6 (glucose) (g/mol):

61

Molar mass of C12H22O11 (sucrose) (g/mol):

62

Molar mass of HCl (g/mol):

63

Molar mass of HNO3 (g/mol):

64

Molar mass of H2SO4 (g/mol):

58

46

57

58

60

180

59

60

62

36.5

61

342

64

98

63

63

65

Molar mass of H3PO4 (g/mol):

66

Molar mass of NaCl (g/mol):

67

Molar mass of KCl (g/mol):

68

Molar mass of NaOH (g/mol):

69

Molar mass of KOH (g/mol):

70

Molar mass of Ca(OH)2 (g/mol):

71

Molar mass of Mg(OH)2 (g/mol):

72

Molar mass of CaCO3 (g/mol):

66

58.5

65

98

68

40

67

74.5

70

74

69

56

72

100

71

58

73

Molar mass of NaHCO3 (g/mol):

74

Molar mass of Na2CO3 (g/mol):

75

Molar mass of MgCO3 (g/mol):

76

Molar mass of NaNO3 (g/mol):

77

Molar mass of KNO3 (g/mol):

78

Molar mass of NH4Cl (g/mol):

79

Molar mass of NH4NO3 (g/mol):

80

Molar mass of (NH4)2SO4 (g/mol):

74

106

73

84

76

85

75

84

78

53.5

77

101

80

132

79

80

81

Molar mass of Na2SO4 (g/mol):

82

Molar mass of MgSO4 (g/mol):

83

Molar mass of CaCl2 (g/mol):

84

Molar mass of Ca(NO3)2 (g/mol):

85

Molar mass of SO2 (g/mol):

86

Molar mass of SO3 (g/mol):

87

Molar mass of NO2 (g/mol):

88

Molar mass of N2O (g/mol):

82

120

81

142

84

164

83

111

86

80

85

64

88

44

87

46

89

Molar mass of O3 (g/mol):

90

Molar mass of Al2O3 (g/mol):

91

Molar mass of Fe2O3 (g/mol):

92

Molar mass of FeO (g/mol):

93

Molar mass of CaO (g/mol):

94

Molar mass of CuSO4 (g/mol):

95

Molar mass of CuSO4 · 5H2O (g/mol):

96

Molar mass of AgNO3 (g/mol):

90

102

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89

48

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92

72

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91

160

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94

159.5

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93

56

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96

170

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95

249.5

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97

Molar mass of $BaSO_4$ (g/mol):

98

Quick trap: CO and N_2 have the same molar mass. What is it?

99

Another sneaky one: $NaHCO_3$ and $MgCO_3$ have the same molar mass. What is it?

100

H_2SO_4 and H_3PO_4 both have molar mass...?

101

9 g of water is how many moles? (H_2O = 18 g/mol)

102

How many grams are in 0.25 mol of CO_2? (44 g/mol)

103

How many moles are in 22 g of CO_2? (44 g/mol)

104

How many moles are in 58.5 g of NaCl?

98

28 g/mol.

97

233

100

98 g/mol.

99

84 g/mol.

102

11 g.

101

0.50 mol.

104

1.00 mol.

103

0.50 mol.

105

How many molecules are in 0.50 mol of anything?

106

How many moles is 1.204×10^{24} molecules?

107

2H2 + O2 -> 2H2O. If you have 3 mol O2 (and lots of H2), max moles H2O you can make?

108

2H2 + O2 -> 2H2O. If you have 5 mol H2 (and lots of O2), max moles H2O?

109

N2 + 3H2 -> 2NH3. If you have 3 mol H2 (excess N2), how many mol NH3?

110

N2 + 3H2 -> 2NH3. If you have 1 mol N2 (excess H2), how many mol NH3?

111

CH4 + 2O2 -> CO2 + 2H2O. If you burn 1 mol CH4, moles of CO2 produced?

112

CH4 + 2O2 -> CO2 + 2H2O. If you burn 2 mol CH4 (excess O2), moles of H2O produced?

106

2.00 mol.

105

3.01×10^{23} particles.

108

5 mol H2O.

107

6 mol H2O.

110

2 mol NH3.

109

2 mol NH3.

112

4 mol H2O.

111

1 mol CO2.

113

CaCO3 -> CaO + CO2. If 0.20 mol CaCO3 decomposes, how many mol CO2?

114

CaCO3 -> CaO + CO2. If you start with 10.0 g CaCO3 (100 g/mol), moles CO2 produced?

115

HCl + NaOH -> NaCl + H2O. If you have 0.30 mol HCl and 0.30 mol NaOH, which is limiting?

116

Limiting reagent in one sentence: it's the reactant that...

117

Fast limiting reagent method that avoids confusion:

118

2H2 + O2 -> 2H2O. You have 4 mol H2 and 1 mol O2. Limiting reactant?

119

Same reaction: 4 mol H2 + 1 mol O2 -> how many mol H2O form?

120

Same reaction: how many mol H2 are left over after reaction?

114

0.10 mol CO_2.

113

0.20 mol CO_2.

116

Runs out first (based on the balanced mole ratio).

115

Neither: it's exactly stoichiometric (1:1).

118

O_2 is limiting.

117

Convert each reactant to moles of the SAME product. Smaller product amount = limiting reactant.

120

2 mol H_2 left.

119

2 mol H_2O.

121

N2 + 3H2 -> 2NH3. You have 2 mol N2 and 3 mol H2. Limiting reactant?

122

Same reaction: with 2 mol N2 and 3 mol H2, how many mol NH3 can form?

123

Same reaction: how many mol N2 are left?

124

CH4 + 2O2 -> CO2 + 2H2O. You have 1.0 mol CH4 and 1.5 mol O2. Limiting reactant?

125

Same reaction: with 1.0 mol CH4 and 1.5 mol O2, how many mol CO2 form?

126

Trap: to find limiting reactant, compare grams of reactants directly. True or false?

127

Theoretical yield means...

128

Actual yield means...

122

2 mol NH3.

121

H2 is limiting.

124

O2 is limiting.

123

1 mol N2 left.

126

False.

125

0.75 mol CO2.

128

The product you ACTUALLY got in the lab.

127

The maximum product you CAN make if everything reacts perfectly.

129

Percent yield formula:

130

Trap: percent yield can be above 100% and that's always fine. True or false?

131

Percent yield = ({{c1::actual}}/{{c2::theoretical}}) x 100%.

132

Molarity (M) tells you...

133

If you have 250 mL of a 0.20 M NaCl solution, how many moles of NaCl is that?

134

Same question but trap version: 250 mL of 0.20 M has moles =

135

How many grams of NaCl to make 500 mL of 0.10 M NaCl? (NaCl=58.5 g/mol)

136

Dilution equation you should know:

130

False (usually).

129

percent yield = (actual / theoretical) x 100%.

132

How many moles of solute are in 1 liter of solution.

131

Percent yield = (actual/theoretical) x 100%.

134

0.050 mol, not 50 mol.

133

0.050 mol.

136

M1V1 = M2V2 (same solute, just adding solvent).

135

2.93 g (about 2.9 g).

137

2.0 M stock -> make 250 mL of 0.50 M. How many mL of stock do you need?

138

Trap: in M1V1=M2V2, V is the volume of solvent. True or false?

139

If you dilute a solution by adding water, what happens to moles of solute?

140

Quick check: 5.85 g NaCl per liter corresponds to what molarity? (NaCl=58.5 g/mol)

141

Concentration trap: if a label says 10% (w/w), that means...

142

ppm (parts per million) in water is often treated as...

143

Dilution: {{c1::M1V1}} = {{c2::M2V2}}.

144

Ideal gas law:

138

False.

137

62.5 mL.

140

0.10 M.

139

They stay the same.

142

mg/L (because 1 L water ~ 1 kg).

141

10 g solute per 100 g solution.

144

$PV = nRT$.

143

Dilution: $M_1V_1 = M_2V_2$.

145

Trap: in PV=nRT, temperature must be in...

146

Celsius to Kelvin conversion:

147

At 0°C and 1 atm (classic STP), 1 mol of ideal gas occupies about...

148

At 25°C and 1 atm (room-ish), 1 mol of gas is closer to...

149

If a gas volume is given at STP and you need moles fast:

150

Example: 11.2 L of O2 at STP is how many moles?

151

Example: 44.8 L of CO2 at STP is how many moles?

152

If you only change temperature (P constant), gas volume is proportional to...

146

$K = °C + 273$.

145

Kelvin (K).

148

24 L (about 24.5 L).

147

22.4 L.

150

0.50 mol.

149

$n = V / 22.4$ (with V in liters).

152

Temperature in Kelvin.

151

2.00 mol.

153

R value you see most with L and atm units is about...

154

Ideal gas law: {{c1::PV}} = {{c2::nRT}}.

155

At STP, 1 mol ideal gas volume (L/mol):

156

Common polyatomic ion: Ammonium

157

Common polyatomic ion: Hydroxide

158

Common polyatomic ion: Nitrate

159

Common polyatomic ion: Nitrite

160

Common polyatomic ion: Sulfate

154

Ideal gas law: $PV = nRT$.

153

$0.0821\ \text{L} \cdot \text{atm}/(\text{mol} \cdot \text{K})$.

156

$NH4^{+}$ (charge $+1$)

155

22.4

158

$NO3^{-}$ (charge -1)

157

OH^{-} (charge -1)

160

$SO4^{2-}$ (charge -2)

159

$NO2^{-}$ (charge -1)

161

Common polyatomic ion: Sulfite

162

Common polyatomic ion: Carbonate

163

Common polyatomic ion:
Bicarbonate (hydrogen carbonate)

164

Common polyatomic ion: Phosphate

165

Common polyatomic ion: Acetate

166

Common polyatomic ion: Permanganate

167

Fast way to write an ionic formula:

168

Trap: ionic formula keeps the charges written in it (like $Na^{+}Cl^{-}$). True or false?

162

CO_3^{2-} (charge -2)

161

SO_3^{2-} (charge -2)

164

PO_4^{3-} (charge -3)

163

HCO_3^- (charge -1)

166

MnO_4^- (charge -1)

165

CH_3COO^- (charge -1)

168

False.

167

Make total charge = 0 (use the smallest whole-number ratio).

169

When do you need parentheses in a formula?

170

Trap: Ca(NO3)2 can be written as CaNO32. True or false?

171

Write the formula for calcium chloride.

172

Write the formula for aluminum oxide.

173

Write the formula for magnesium hydroxide.

174

Write the formula for sodium carbonate.

175

Write the formula for calcium nitrate.

176

Write the formula for ammonium sulfate.

170

False.

169

When you have more than one of a polyatomic ion.

172

Al_2O_3

171

$CaCl_2$

174

Na_2CO_3

173

$Mg(OH)_2$

176

$(NH_4)_2SO_4$

175

$Ca(NO_3)_2$

177

Write the formula for aluminum sulfate.

178

Write the formula for sodium phosphate.

179

Write the formula for ammonium nitrate.

180

Iron and copper trap: why do they often have roman numerals in names (iron(II), iron(III))?

181

Example: iron(III) chloride formula is...

182

Example: iron(II) chloride formula is...

183

Parentheses rule: use parentheses when a {{c1::polyatomic ion}} appears more than {{c2::once}} in the formula (ex: Ca(NO3)2).

184

Percent composition (by mass) of an element in a compound is...

178

Na3PO4

177

Al2(SO4)3

180

Because they can form more than one ion charge.

179

NH4NO3

182

FeCl2.

181

FeCl3.

184

(mass of that element in 1 mole of compound / molar mass of compound) x 100%.

183

Parentheses rule: use parentheses when a polyatomic ion appears more than once in the formula (ex: Ca(NO3)2).

185

Empirical formula means...

186

Molecular formula means...

187

Trap: empirical formula and molecular formula are always the same. True or false?

188

Fast empirical formula workflow from percentages:

189

If your mole ratios come out like 1 : 1.5 : 1, what do you do?

190

From empirical formula to molecular formula: key step is...

191

Example idea: empirical formula CH_2O has mass 30 g/mol. If molar mass is 180, the multiplier is...

192

Combustion analysis core idea (high-level):

186

Actual numbers of each atom in the molecule.

185

Simplest whole-number ratio of atoms in a compound.

188

1) Assume 100 g. 2) Convert g to moles. 3) Divide by smallest. 4) Multiply to clear decimals.

187

False.

190

Find the multiplier: (molar mass / empirical formula mass).

189

Multiply everything by 2 -> 2 : 3 : 2.

192

All C ends up in CO2, all H ends up in H2O (use masses to back-calculate C and H in the original compound).

191

6 (and molecular formula is C6H12O6).

193

Empirical formula = {{c1::simplest whole-number ratio}} of atoms; molecular formula = {{c2::actual}} numbers of atoms.

193

Empirical formula = simplest whole-number ratio of atoms; molecular formula = actual numbers of atoms.