

There are $5.01(10^{22})$ carbon atoms in the diamond.

6. A bottle containing 100 tablets contains 200 mg of copper.

Molar mass of Cu = 63.55 g/mol

$$\frac{0.200 \text{ g}}{63.55 \text{ g/mol}} = 3.15 \times 10^{-3} \text{ mol of Cu}$$

If there is 3.15×10^{-3} mol of Cu in the bottle, then there must be the same number of moles of CuO present.

Molar mass of CuO = 79.56 g/mol

Mass of CuO in the bottle = $(3.15 \times 10^{-3} \text{ mol})(79.56 \text{ g/mol}) = 0.2506 \text{ g}$
or 250.6 mg of CuO.

$$\text{Mass of CuO in one tablet} = \frac{250.6 \text{ mg}}{100 \text{ tablets}} = 2.5 \text{ mg CuO per tablet.}$$

Chapter 5 Review Answers

Student Book pages 193–195

Answers to Knowledge/Understanding Questions

- Atomic mass is the mass of one atom of a specific isotope. Average atomic mass is the weighted average of atomic masses for the naturally occurring mix of isotopes of an element. For example, chlorine-35 has atomic mass 34.97 u while the average atomic mass for chlorine is 35.45 u (since there is also chlorine-37 to be considered).
- The average has to be weighted to reflect the relative abundances of the various isotopes, which are not present in equal proportions.
- Chlorine has two isotopes, chlorine-35 and chlorine-37. The average atomic mass of chlorine is due to the weighted average of the two isotopes, therefore, no one chlorine atom has a mass of 35.45 u.
- The Avogadro constant provides a conversion between the mass of one mole of an entity and the average atomic mass of that same entity. N_A atoms with an average atomic mass of x u will have a mass of x g.
- A chemist can measure the mass of samples using a balance. Because the chemist knows the molar mass of substances, they can use this knowledge to calculate the number of particles in that sample of substance.
- (a) The mole is defined as an amount of a substance that has the same number of particles as exactly 12 g of carbon-12. This number is called the Avogadro constant, and it has been experimentally determined to be 6.02×10^{23} .
(b) convenience
- The average atomic mass of an element, expressed in atomic mass units, has the same numerical value as the molar mass of an element, expressed in grams.
- (a) For a metallic element, the molar mass is the mass of N_A atoms of that element.
(b) For a diatomic element, the molar mass is equal to the mass of $2 \times N_A$ atoms of the element.
(c) For a compound, the molar mass is equal to the combined mass of N_A atoms for each atom found in the compound.

Answers to Inquiry Questions

9.

Isotope	Percentage Abundance	Approximate Atomic Mass (u)
Ar-36	0.34	36
Ar-38	0.06	38
Ar-40	99.6	40

$$\text{average atomic mass} = \frac{[0.34(36.0) + 0.06(38.0) + 99.6(40)] \text{ u}}{100} = \frac{[12.24 + 2.28 + 3984] \text{ u}}{100} = 40 \text{ u}$$

10.

Isotope	Percentage Abundance	Atomic Mass (u)
Ga-69	60.0	69
Ga-71	40.0	71

$$\text{average atomic mass} = \frac{[60(69.0) + 40(71.0)] \text{ u}}{100} = \frac{[4140 + 2840] \text{ u}}{100} = 70 \text{ u}$$

11.

Isotope	Percentage Abundance	Atomic Mass (u)
Ge-70	20.5	70
Ge-72	27.4	72
Ge-73	7.8	73
Ge-74	36.5	74
Ge-76	7.8	76

$$\text{average atomic mass} = \frac{[(20.5)(70) + (27.4)(72) + (7.8)(73) + (36.5)(74) + (7.8)(76)] \text{ u}}{100}$$

$$\text{average atomic mass} = \frac{[1435 + 1972.8 + 569.4 + 2701 + 592.8] \text{ u}}{100} = 72.7 \text{ u}$$

12.

Isotope	Percentage Abundance	Atomic Mass (u)
K-39	x	39
K-41	$100 - x$	41

Given that the average atomic mass = 39.1 u

$$39.1 \text{ u} = \frac{[x(39.0) + (100 - x)(41.0)] \text{ u}}{100}$$

$$3910 = [39.0x + 4100 - 41.0x]$$

$$2.0x = 190 \text{ so } x = 95.$$

Thus natural potassium is 95% K-39 and 5% K-41.

13. In each case, divide the mass by the molar mass ($n = \frac{m}{M}$).

$$(a) n = \frac{0.453 \text{ g}}{159.7 \text{ g/mol}} = (2.84)(10^{-3}) \text{ mol}$$

$$(b) n = \frac{50.7 \text{ g}}{98.09 \text{ g/mol}} = 0.517 \text{ mol}$$

$$(c) n = \frac{(1.24)(10^{-2}) \text{ g}}{152.00 \text{ g/mol}} = (8.15)(10^{-5}) \text{ mol}$$

$$(d) n = \frac{(8.2)(10^2) \text{ g}}{187.37 \text{ g/mol}} = 4.38 \text{ mol}$$

$$(e) n = \frac{12.3 \text{ g}}{97.95 \text{ g/mol}} = 0.126 \text{ mol}$$

14. In each case, divide by the Avogadro constant ($n = \frac{N}{N_A}$).

$$(a) n = \frac{(4.27)(10^{21})}{(6.02)(10^{23}) \text{ mol}^{-1}} = (7.09)(10^{-3}) \text{ mol}$$

$$(b) n = \frac{(7.39)(10^{23})}{(6.02)(10^{23}) \text{ mol}^{-1}} = 1.23 \text{ mol}$$

$$(c) n = \frac{(5.38)(10^{22})}{(6.02)(10^{23}) \text{ mol}^{-1}} = (8.94)(10^{-2}) \text{ mol}$$

$$(d) n = \frac{(2.91)(10^{23})}{(6.02)(10^{23}) \text{ mol}^{-1}} = 0.483 \text{ mol}$$

$$(e) n = \frac{(1.62)(10^{24})}{(6.02)(10^{23})\text{mol}^{-1}} = 2.69 \text{ mol}$$

$$(f) n = \frac{(5.58)(10^{20})}{(6.02)(10^{23})\text{mol}^{-1}} = (9.27)(10^{-4}) \text{ mol}$$

15.

Isotope	Molar Mass (g/mol)	Sample Mass (g)	Number of Molecules	Number of Moles of Molecules	Number of Moles of Atoms
NaCl	58.44	58.44	$(6.02)(10^{23})$	1.00	2.00
NH ₃	17.04	24.8	$(8.79)(10^{23})$	1.46	5.84
H ₂ O	18.02	1.58	$(5.28)(10^{22})$	$(8.77)(10^{-2})$	$(2.63)(10^{-1})$
Mn ₂ O ₃	157.88	10.5	$(4.00)(10^{22})$	$(6.64)(10^{-2})$	0.332
K ₂ CrO ₄	194.20	$(9.67)(10^{-1})$	$(3.00)(10^{21})$	$(4.98)(10^{-3})$	$(3.49)(10^{-2})$
C ₈ H ₈ O ₃	152.16	$(1.99)(10^3)$	$(7.90)(10^{24})$	13.1	249
Al(OH) ₃	78.01	$(6.66)(10^4)$	$(5.14)(10^{26})$	$(8.54)(10^2)$	$(5.98)(10^3)$

16. (a) PtBr₂: $M = 195.08 \text{ g/mol} + 2(79.90) \text{ g/mol} = 354.88 \text{ g/mol}$

(b) C₃H₅O₂H:

$$M = 3(12.01) \text{ g/mol} + 5(1.01) \text{ g/mol} + 2(16.00) \text{ g/mol} + 1.01 \text{ g/mol}$$

$$= 74.09 \text{ g/mol}$$

(c) Na₂SO₄: $M = 2(22.99) \text{ g/mol} + 32.07 \text{ g/mol} + 4(16.00) \text{ g/mol} = 142.05 \text{ g/mol}$

(d) (NH₄)₂Cr₂O₇:

$$M = 2[14.01 \text{ g/mol} + 4(1.01) \text{ g/mol}] + 2(52.00) \text{ g/mol} + 7(16.00) \text{ g/mol}$$

$$= 252.10 \text{ g/mol}$$

(e) Ca₃(PO₄)₂: $M = 3(40.08) \text{ g/mol} + 2[30.97 \text{ g/mol} + 4(16.00) \text{ g/mol}]$
 $= 310.18 \text{ g/mol}$

(f) Cl₂O₇: $M = 2(35.45) \text{ g/mol} + 7(16.00) \text{ g/mol} = 182.90 \text{ g/mol}$

17. (a) Sample = 3.70 mol = x g

Mole = 1.0 mol = 18.02 g

$$\frac{3.70 \text{ mol}}{1.0 \text{ mol}} = \frac{x}{18.02 \text{ g}}$$

$$x = 66.7 \text{ g}$$

There are 66.7 g of water.

1 mole H ₂ O $(6.02)(10^{23})$ molecules 18.02 g

(b) Sample = $(8.43)(10^{23})$ molecules = x g

Mole = $(6.02)(10^{23})$ molecules = 239.20 g

$$\frac{(8.43)(10^{23}) \text{ molecules}}{(6.02)(10^{23}) \text{ molecules}} = \frac{x}{239.20 \text{ g}}$$

$$x = 335 \text{ g}$$

There are 335 g in the sample.

1 mole PbO ₂ $(6.02)(10^{23})$ molecules 239.20 g
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(c) Sample = 14.8 mol = x g

Mole = 1.0 mol = 253.33 g

$$\frac{14.8 \text{ mol}}{1.0 \text{ mol}} = \frac{x}{253.33 \text{ g}}$$

$$x = (3.75)(10^3) \text{ g}$$

There are $(3.75)(10^3)$ g in the sample.

1 mole BaCrO ₄ $(6.02)(10^{23})$ molecules 253.33 g
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(d) Sample = $(1.23)(10^{22})$ molecules = x g

Mole = $(6.02)(10^{23})$ molecules = 70.90 g

$$\frac{(1.23)(10^{22}) \text{ molecules}}{(6.02)(10^{23}) \text{ molecules}} = \frac{x}{70.90 \text{ g}}$$

$$x = 1.45 \text{ g}$$

There are 1.45 g in the sample.

1 mole Cl ₂ $(6.02)(10^{23})$ molecules 70.90 g
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(e) Sample = $(9.48)(10^{23})$ molecules = x g

Mole = $(6.02)(10^{23})$ molecules = 36.46 g

$$\frac{(9.48)(10^{23}) \text{ molecules}}{(6.02)(10^{23}) \text{ molecules}} = \frac{x}{36.46 \text{ g}}$$

$$x = 57.4 \text{ g}$$

1 mole HCl $(6.02)(10^{23})$ molecules 36.46 g
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There are 57.4 g in the sample.

$$\begin{aligned} \text{(f) Sample} &= (7.74)(10^{19}) \text{ molecules} = x \text{ g} \\ \text{Mole} &= (6.02)(10^{23}) \text{ molecules} = 159.70 \text{ g} \\ \frac{(7.74)(10^{19}) \text{ molecules}}{(6.02)(10^{23}) \text{ molecules}} &= \frac{x}{159.70 \text{ g}} \end{aligned}$$

$$x = (2.05)(10^{-2}) \text{ g}$$

There are $(2.05)(10^{-2})$ g in the sample.

$$\begin{aligned} &1 \text{ mole Fe}_2\text{O}_3 \\ &(6.02)(10^{23}) \text{ molecules} \\ &159.70 \text{ g} \end{aligned}$$

18. C_6H_6 : $M = 78.12 \text{ g/mol}$ and given $m = 45.6 \text{ g}$

$$n = \frac{m}{M} = \frac{45.6 \text{ g}}{78.12 \text{ g/mol}} = 0.584 \text{ mol}$$

Using $n = \frac{N}{N_A}$, the number of molecules is

$$N = (0.584 \text{ mol})(6.02)(10^{23}) \text{ molecules/mol} = (3.52)(10^{23}) \text{ molecules.}$$

19. Since there are 4 atoms per molecule, there are:

$$4(0.72 \text{ mol})(6.02 \times 10^{23} \text{ atoms/mol}) = (4.22)(10^{24}) \text{ atoms.}$$

20. (a) for one atom, mass = 131.29 u.

(b) for one mole, mass = 131.29 g.

$$\text{(c) } 131.29 \text{ g} = (6.02)(10^{23})(131.29 \text{ u}) \text{ or } 1 \text{ g} = (6.02)(10^{23}) \text{ u}$$

$$\text{so } 1 \text{ u} = \left[\frac{1.00}{(6.02)(10^{23})} \right] \text{ g} = (1.66)(10^{-24}) \text{ g.}$$

$$\text{(d) Mass of one atom is } 131.29 \text{ u} = 131.29(1.66)(10^{-24}) \text{ g} = (2.18)(10^{-22}) \text{ g.}$$

$$\text{(e) Mass of one mole is } 131.29 \text{ g} = 131.29(6.02)(10^{23}) \text{ u} = (7.90)(10^{25}) \text{ u.}$$

21. First find the number of moles of C atoms by addition:

$$n = 1(0.237 \text{ mol}) + 2(2.38 \text{ mol}) = 5.00 \text{ mol}; \text{ then the number of atoms is:}$$

$$5.00 \text{ mol} (6.02)(10^{23} \text{ atoms/mol}) = (3.01)(10^{24} \text{ atoms})$$

22. Consider the water first. In $(3.49)(10^{23})$ molecules there are $2(3.29)(10^{23})$ or $(6.98)(10^{23})$ H atoms.

$$\text{Now consider CH}_3\text{OH. Using } n = \frac{m}{M}, \text{ the number of moles of molecules is}$$

$$\frac{78.1 \text{ g}}{32.05 \text{ g/mol}}, \text{ or } 2.44 \text{ moles.}$$

Since each molecule contains four atoms of H, there are $4(2.44 \text{ mol})$ or 9.76 moles of H atoms.

$$\text{This is } 9.76 \text{ mol}(6.02)(10^{23} \text{ atoms/mol}) \text{ or } (5.88)(10^{24}) \text{ atoms.}$$

The total number of H atoms is

$$(6.98)(10^{23}) \text{ atoms} + (5.88)(10^{24}) \text{ atoms} = (6.58)(10^{24}) \text{ atoms.}$$

23. There are two nitrate ions per formula unit of the compound.

$$\text{Thus there are } 2(3.76)(10^{-1}) \text{ mol} = (7.52)(10^{-1}) \text{ mol nitrate ions.}$$

$$\text{Thus there are } (7.52)(10^{-1} \text{ mol})(6.02)(10^{23} \text{ ions/mol}) = (4.53)(10^{23}) \text{ nitrate ions.}$$

24. The required mole ratio is $\frac{\text{oxygen}}{\text{ethanol}} = \frac{3}{1}$. Find the number of moles of ethanol.

$$n = \frac{m}{M} = \frac{92.0 \text{ g}}{46.08 \text{ g/mol}} = 2 \text{ mol. Thus } 3(2) \text{ mol} = 6 \text{ mol of oxygen molecules are required.}$$

Thus $6(32.00) \text{ g}$ or 192 g of oxygen are required.

25.

Isotope	Percentage Abundance	Atomic Mass (u)
Br-79	x	79.0
Br-81	$100 - x$	81.0

From the periodic table, average atomic mass = 79.9 u.

$$79.9 \text{ u} = \frac{[x(79.0) + (100 - x)(81.0)] \text{ u}}{100}$$

$$7990 \text{ u} = [79.0x + 8100 - 81.0x] \text{ u}$$

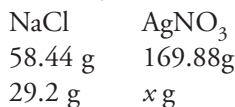
$$2.0x = 110, \text{ so } x = 55.0$$

Natural bromine is 55.0% Br-79 and 45.0% Br-81.



One formula unit reacts with one formula unit, so one mole of sodium chloride reacts with one mole of silver nitrate.

(b) To solve, find the molar masses and use ratio and proportion.



$$\text{Thus } \frac{58.44 \text{ g}}{29.2 \text{ g}} = \frac{169.88 \text{ g}}{x}$$

$$x = 84.9$$

Therefore, 84.9 g of silver nitrate will react with 29.2 g of sodium chloride.

27. The relative masses of atoms must be the same on all planets so we can use the values on our Periodic Table.

Entity	Number of Wogs	Mass in Wibbles
C-12	1.0	12.0
(a) N	1.0	14.01
(b) O ₂	(5.00)(10 ⁻¹)	(5.00)(10 ⁻¹)(32.00) = 16.0

(c) 1 wog = (2.50)(10²¹) entities.

Converting to moles, this is $\frac{(2.50)(10^{21}) \text{ atom/wog}}{(6.02)(10^{23}) \text{ atom/mol}} = (4.15)(10^{-3}) \text{ mol/wog}$,

so 1 mol = (4.15)(10⁻³) wog.

Thus, 1 wog of H atoms has a mass of (4.19)(10⁻³) g.

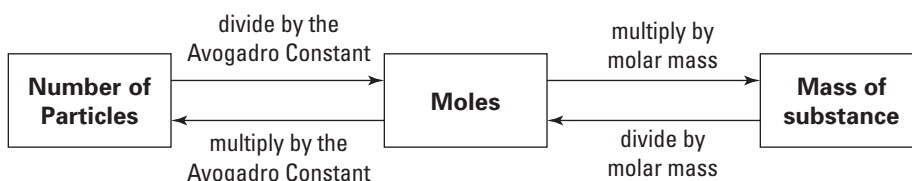
Answers to Communication Questions

28. The Avogadro constant is the number of entities in a mole. A mole is defined as an amount of matter having as many entities as there are atoms in 12 g of carbon-12.

This definition does not tell us what the number of entities is. We must carry out an experiment that allows for the indirect counting of the number of entities in a mole.

29. The atomic mass scale was defined with C-12 having mass 12 u. Since proton and neutron masses are not exactly 1 u, — and electrons have mass — all other masses are not whole numbers.

30.



31. Avogadro's constant is applied to atoms, elements, molecules, compounds, formula units, or ions. N_A describes the relationship between a single particle and 6.02×10^{23} particles. This conversion only applies when moving between the same particle, not different particles. Carbon dioxide is a molecule that contains three atoms.

One mole of carbon dioxide contains 6.02×10^{23} molecules of carbon dioxide, or $(3)(6.02 \times 10^{23}) = 1.8 \times 10^{24}$ atoms.

Answers to Making Connections Questions

32. (a) Molar mass of Fe = 55.85 g/mol

number moles of Fe required per day = $0.0148 \text{ g} \div 55.85 \text{ g/mol}$

= $2.65 \times 10^{-4} \text{ mol Fe}$

(b) There is one mole of Fe atoms per molecule of ferrous gluconate, therefore the number of moles of iron taken per day is equal to the number of moles of ferrous gluconate required per day.

Molar mass of $\text{Fe}(\text{C}_6\text{H}_{11}\text{O}_7)_2 = 446.33 \text{ g/mol}$

Number of moles of $\text{Fe}(\text{C}_6\text{H}_{11}\text{O}_7)_2$ required per day = $2.65 \times 10^{-4} \text{ mol}$

Mass of $\text{Fe}(\text{C}_6\text{H}_{11}\text{O}_7)_2$ required = $(446.33 \text{ g/mol})(2.65 \times 10^{-4} \text{ mol}) = 0.1182 \text{ g}$,
or 118.2 mg

Therefore 118.2 mg of ferrous gluconate is required per day.

(c) For information regarding the addition of elemental iron to cereals, check out Health Canada's web site and Food Guide for more information:

<http://www.hc-sc.gc.ca>.

33. Molar mass of niacin = 123.14 g/mol

Molar mass of vitamin B (niacinamide) = 122.15 g/mol.

(a) Number of moles of niacinamide in a vitamin supplement
= $0.100 \text{ g} \div 122.15 \text{ g/mol} = 8.19 \times 10^{-4} \text{ mol}$

Equivalent mass of niacin = $(8.19 \times 10^{-4} \text{ mol})(123.14 \text{ g/mol})$
= 0.10085 g = 100.9 mg

(b) The recommended nutrient intake (RNI) for niacin is:

average adult male: 14 – 20 mg

average adult female: 10 – 14 mg

(c) Foods high in niacin include tuna, wheat bran, and Ovaltine.

(d) Symptoms of a niacin overdose include: blurred vision, excessive itching, nausea, abdominal pain, and lightheadedness. Overdose is usually not life threatening but in any case, the nearest poison control center should be called.

(e) More details on dietary inputs and vitamins can be found at the following sites:

<http://webmd.lycos.com> and

<http://healthyeating.org/books/foodfacts/html/data/data-fs.html>.