Molecular mass and moles worksheet answers

a) We have 1/2 a mole $Ca(IO_3)_2$. There are two iodine atoms per molecule. Step #1 molecules = 0.50 mol X 6.02 X 1023 molecules/mole = 3.01 X 1023 molecules of I Step #2 in atoms I = molecules X 2 molecule atoms I = 3.01 X 1023 molecules of atoms I X 2 molecule I = 6.02 X 10 10 23 Atoms of I in 0.5 mol de $Ca(IO_3)_2$ b) m = n • M = 0.5 mol X 389.88 g/mol = 194.9 4 grams of $Ca(IO_3)_2$ In Dalton's theory each chemical compound has a particular combination of atoms and that the proportions of the numbers of atoms of the present elements are usually small integers. We have also described the multi-proportion law, which states that the ratios of the masses of elements that make up a number of compounds are small whole numbers. The problem for Dalton and other early chemists was to discover the quantitative relationship between the number of atoms in a chemical and its mass. Because the masses of individual atoms are so tiny (in order of 10−23 g/atom), chemists do not measure the mass of individual atoms or molecules. In the laboratory, for example, masses of compounds and elements used by chemists typically range from milligrams to grams, while in industry, chemicals are bought and sold in kilograms and tons. To analyse the transformations that occur between individual atoms or molecules in a chemical reaction, it is therefore absolutely imperative that chemists know how many atoms or molecules are contained in a measurable amount in the laboratory, a certain mass of sample. The unit that provides this link is the mole. The amount of a substance containing the same number of units (e.g. atoms or molecules) as the number of carbon atoms in exactly 12 g of isotopically pure carbon-12., from Latin moles, meaning stack or heap (not from the small underground animal!). Many family items are sold in numerical quantities that have unusual names. For example, soda ments come in a six packet, eggs are sold by the dozen (12), and pencils often come in a brute (12 dozen, or 144). Printer sheets are packed in 500-binned disks, a seemingly large number. Atoms are so small, however, that even 500 atoms are too small to see or measure by the most common techniques. Any easily measurable mass of an element or compound contains an extraordinarily large number of atoms, molecules or ions, so an extraordinarily large numerical unit is needed to count them. The mole is used for this purpose. A mole is defined as the of a substance containing the number of carbon atoms in exactly 12 g of isotopically pure carbon-12. According to the latest experimental measurements, this carbon mass-12 contains 6.022142 × 1023 atoms, but for most purposes 6,022 × 1023 provides an adequate number of significant figures. Just as 1 mol of atoms contains 6,022 atoms × 1023, 1 egg mol contains 6,022 × 1023 eggs. The number on a mole is called the Avogadro number: 6.022142 x 1023, after the 19th-century Italian scientist who first proposed how to measure the number of molecules in a gas. Since the mass of gas can also be measured in a sensitive balance, knowing both the number of molecules and their total mass allows us to simply determine the mass of a single molecule in grams. The mole provides a bridge between the atomic world (amu) and the laboratory (grams). It allows the determination of the number of molecules or atoms weighing them. The numerical value of the Avogadro number, usually written as No, is a consequence of the arbitrary value of a kilogram, a Pt-Ir metal block called the International Kilogram Prototype, and the choice of reference for the atomic mass unit scale, a carbon-12 atom. A mole of C-12 by definition weighs exactly 12 g and the number of Avogadro is determined by counting the number of atoms. It's not that easy. The number of Avogadro is the fundamental constant that is determined less accurately. The definition of a mole —i.e. the decision to base it on 12 g of carbon-12— is arbitrary, but one came after a discussion between chemists and physicists discussing whether to use natural carbon, a mixture of C-12 and C-13, or hydrogen. The important point is that 1 carbon mole –or anything else, whether atoms, compact discs or houses- always has the same number of objects: 6,022 × 1023. In the following video, Prof Steve Boon shows how the Avogadro hypothesis can be used to measure the molecular masses of He, N_2 and CO_2. Go ahead and record the measurements to get the relative masses. When we consider the behavior of gases in Unit 5, we can use the data to calculate the molecular weight of each gas. This method was, until the invention of the mass spectrometer, the best way to measure the molecular weights of gas molecules A mole always has the same number of objects: 6,022 × 1023. To appreciate the magnitude of Avogadro's number, consider a penny mole. Stacked vertically, a penny mole would be 4.5 × 1017 me tall, or nearly six times the diameter of the Milky Way galaxy. If a penny mole was distributed equally among the entire population on Earth, each person would get more than a trillion dollars. Clearly, the mole is so large that it is only useful for measuring very small objects, such as atoms. The concept of the mole count a specific number of measurable amounts of elements and compounds. To obtain 1 mol of carbon-12 atoms, we would weigh 12 g of isotopically pure carbon-12. Because each element has a different atomic mass, however, a mole in each element has a different mass, although it contains the same number of atoms (6,022 × 1023). This is analogous to the fact that a dozen extra large eggs weigh more than a dozen small eggs, or that the total weight of 50 adult humans is greater than the total weight of 50 children. Because of the way the mole is defined, for each element the number of grams in a mole is the same as the number of atomic mass units in the atomic mass of the element. For example, the mass of 1 magnesium mole (atomic mass = 24,305mu) is 24,305 g. Due to the atomic mass of magnesium (24,305 amu) is slightly more than twice that of a carbon-12 atom (12 amu), the mass of 1 mol of magnesium atoms (24,305 g) is slightly more than twice that of 1 carbon-12 mol (12 g). Similarly, the mass of 1 mole of helium (atomic mass = 4.002602 amu) is 4.002602 g, which is about a third of that of 1 carbon-12 mole. Using the mole concept, we can now restore Dalton's theory: 1 mol of a compound consists of combined elements in quantities because mole proportions are small integers. For example, 1 molecule of water (H_2O) has 2 moles of hydrogen atoms and 1 mol of oxygen atoms. a) We have 1/2 a mole $Ca(IO_3)_2$. There are two iodine atoms per molecule. 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