

Stoichiometry is one of the most important topics present in Advanced Placement Chemistry. It is present within all of the units of study throughout the course, so it is imperative that you are efficient and proficient. In all your calculations you will need to **show your work** explicitly AND always **show units and what substance those units are assigned to**. It is usually safe to round your answers to three significant digits!

The four calculations that you have to master are the following:

- Primarily Solids:

$$\text{grams} \div \text{Molar Mass} = \text{moles} \quad \text{or} \quad \text{Molar Mass} = \frac{\text{grams}}{\text{mole}}$$

- Solutions:

$$\text{Molarity} \times \text{Liters} = \text{moles} \quad \text{or} \quad \text{Molarity} = \frac{\text{moles}}{\text{Liters}}$$

- Gases:

$$PV = nRT \quad \text{or} \quad (P \text{ in atm})(Vol \text{ in L}) = (\text{moles})(0.08206 \frac{\text{L atm}}{\text{mol K}})(Temp \text{ in K})$$

- Converting:

$$\text{molar ratios} \quad \text{or} \quad \text{moles of one substance} : \text{moles of another substance}$$

Here is an example of what your work would look like for solids:

Mass of KI tablet	0.425 g
Mass of thoroughly dried filter paper	1.462 g
Mass of filter paper + precipitate after first drying	1.775 g
Mass of filter paper + precipitate after second drying	1.699 g
Mass of filter paper + precipitate after third drying	1.698 g

A student is given the task of determining the mass percentage of I^- content of tablets that contain KI and an inert, water-soluble sugar as a filler. A tablet is dissolved in 50.0 mL of distilled water, and an excess of 0.20 M $Pb(NO_3)_2(aq)$ is added to the solution. A yellow precipitate of $PbI_2(s)$ forms, which is then filtered, washed, and dried. The data from the experiment are shown in the table above.

Step 1: Find the mass of just the precipitate

$$1.698 \text{ g of FP and } PbI_2 - 1.462 \text{ g of FP} = 0.236 \text{ g of } PbI_2$$

Step 2: Convert the mass of the precipitate into moles of the precipitate by dividing by the molar mass

$$0.236 \text{ g of } PbI_2 \div 461.0 \frac{\text{g}}{\text{mol}} = 0.000512 \text{ moles of } PbI_2$$

Step 3: Convert the moles of the precipitate into moles of the substance you want using a molar ratio

$$1 \text{ mole of } PbI_2 : 2 \text{ moles of } I^-$$

$$0.000512 \text{ moles of } PbI_2 : 0.00102 \text{ moles of } I^-$$

Step 4: Convert the moles of the substance you want into grams of the substance you want

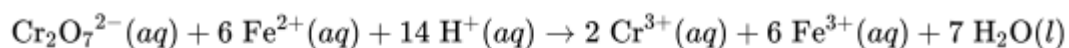
$$0.00102 \text{ moles of } I^- \cdot 126.90 \frac{\text{g}}{\text{mol}} = 0.130 \text{ grams of } I^-$$

Step 5: Find the mass percentage by taking the part divided by the whole

$$\% \text{ mass of } I^- \text{ in tablet} = 0.130 \text{ g of } I^- \div 0.425 \text{ g of tablet} \cdot 100\% = 30.6\% \text{ of } I^-$$

Here is another example of what your work would look like for solutions:

Determine the mass percent of iron in a sample of iron ore by doing a redox titration. The student finds the mass of the sample and then dissolves it in acid to a total volume of 25.00 mL. Using a buret, the student titrates the 25.00 mL solution with 0.017 M $\text{K}_2\text{Cr}_2\text{O}_7$. The following reaction occurs during the titration.



Mass of iron ore in original sample	0.36 g
Buret reading at end point	45.52 mL
Initial buret reading	15.05 mL

Step 1: Find the volume used of the titrant

$$45.52 \text{ mL of } \text{Cr}_2\text{O}_7^{2-} - 15.05 \text{ mL of } \text{Cr}_2\text{O}_7^{2-} = 30.47 \text{ mL of } \text{Cr}_2\text{O}_7^{2-} \text{ used} = 0.03047 \text{ L of } \text{Cr}_2\text{O}_7^{2-} \text{ used}$$

Step 2: Multiply the concentration of titrant by the liters of titrant to find the moles of the titrant

$$0.017 \text{ M } \text{Cr}_2\text{O}_7^{2-} \cdot 0.03047 \text{ L of } \text{Cr}_2\text{O}_7^{2-} = 0.000518 \text{ moles } \text{Cr}_2\text{O}_7^{2-}$$

Step 3: Convert the moles of titrant into moles of the substance you want by using a molar ratio

$$1 \text{ mol of } \text{Cr}_2\text{O}_7^{2-} : 6 \text{ mol of } \text{Fe}^{2+}$$

$$0.000518 \text{ mol of } \text{Cr}_2\text{O}_7^{2-} : 0.00311 \text{ mol of } \text{Fe}^{2+}$$

Step 4: Convert the moles of the substance you want into grams of the substance you want

$$0.00311 \text{ mol of } \text{Fe}^{2+} \cdot 55.85 \frac{\text{g}}{\text{mol}} = 0.173 \text{ g of } \text{Fe}^{2+}$$

Step 5: Find the mass percentage by taking the part divided by the whole

$$\text{Mass \% of Fe in sample} = 0.173 \text{ g of } \text{Fe}^{2+} \div 0.36 \text{ g total} \cdot 100\% = 48.1\% \text{ of iron}$$

Here is another example of a solid that uses a different molar ratio:

Some building materials contain small capsules filled with paraffin wax to improve the insulating properties of the materials. The melting and freezing of paraffin inside the capsules in the material helps to regulate temperature. Paraffin wax begins to melt at 37°C. Calculate the amount of thermal energy, in kJ, that is required to melt 15.2 grams of solid paraffin wax when the temperature of the surroundings is above the melting point of paraffin. (The molar mass of paraffin is 282.62 g/mol, and its molar heat of fusion is 48.78 kJ/mol.)

Step 1: Convert the mass of a substance into moles of that substance by dividing by the molar mass

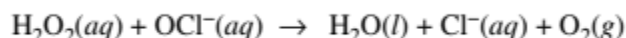
$$15.2 \text{ g of paraffin} \div 282.62 \frac{\text{g}}{\text{mol}} = 0.0538 \text{ mol of paraffin}$$

Step 2: Calculate the heat of the substance using a molar ratio using the molar heat of fusion

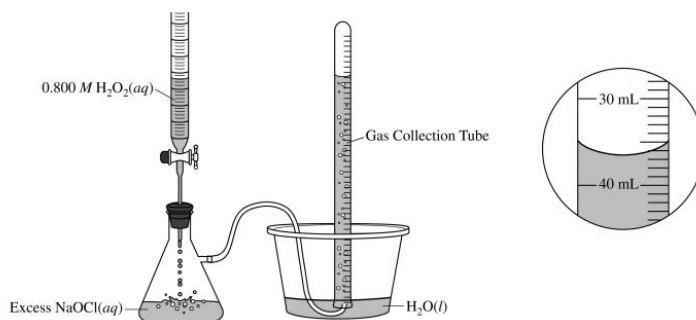
$$48.78 \text{ kJ of heat} : 1 \text{ mol of paraffin}$$

$$? \text{ kJ of heat} : 0.0538 \text{ mol of paraffin}$$

$$2.62 \text{ kJ of heat}$$

Here is one last example of a molar calculation involving gases:

A student investigates the reaction between $\text{H}_2\text{O}_2(aq)$ and $\text{NaOCl}(aq)$, which is represented by the net-ionic equation shown above. The student decides to produce $\text{O}_2(g)$ at a pressure of 0.988 atm and a temperature of 298 K using the reaction represented above. The student uses the equipment shown below. The student sets up a 250 mL Erlenmeyer flask fitted with a one-hole stopper. The flask is connected to a 50 mL gas-collection tube that initially is completely filled with water. Through the chemical reaction, the student produced the volume of gas indicated in the diagram below. The vapor pressure of water at 298 K is 23.8 mmHg.



Step 1: Determine the volume of gas produced to the correct amount of significant digits

$$30.65 \text{ mL}$$

Step 2: Calculate the vapor pressure of water at 298 K in atm using the ratio on the equation sheet

$$760 \text{ mmHg} : 1 \text{ atm}$$

$$23.8 \text{ mmHg} : 0.0313 \text{ atm}$$

Step 3: Determine the gas in the collection tube of just oxygen gas

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2\text{O}}$$

$$0.998 \text{ atm total} = P_{\text{O}_2} + 0.0313 \text{ atm of H}_2\text{O} \quad P_{\text{O}_2} = 0.967 \text{ atm}$$

Step 4: Calculate the moles of oxygen gas produced in the collection tube using $PV = nRT$

$$(0.967 \text{ atm of O}_2)(0.03065 \text{ L of O}_2) = n(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298\text{K})$$

$$n = 0.00121 \text{ moles of O}_2$$

Step 5: Convert the moles of oxygen gas to moles of hydrogen peroxide using a molar ratio

$$1 \text{ mole of O}_2 : 1 \text{ mole of H}_2\text{O}_2$$

$$0.00121 \text{ moles of O}_2 : 0.00121 \text{ moles of H}_2\text{O}_2$$

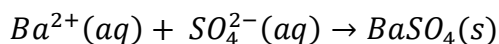
Step 6: Calculate the volume of hydrogen peroxide that was dispensed using its molarity

$$M \times L = \text{moles}$$

$$0.800 \text{ M H}_2\text{O}_2 \times L = 0.00121 \text{ moles of H}_2\text{O}_2$$

$$\text{Volume} = 0.00151 \text{ L} = 1.51 \text{ mL}$$

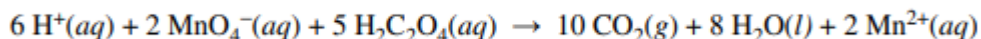
Please note that every step must be shown in detail – including UNITS and the SUBSTANCE those units are assigned to. Just work step by step to get to your answer. The process of the work is as important as the final answer!



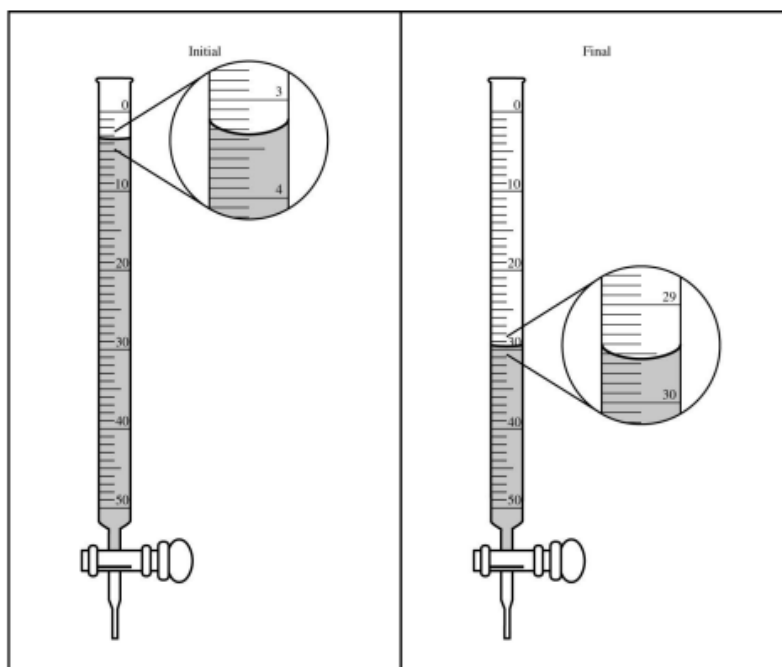
1. A student is given the task of determining the molar concentration of CuSO_4 solution using the precipitation reaction shown above. The student adds 20.0 mL of 0.200 M $\text{Ba}(\text{NO}_3)_2$ to 50.0 mL of the CuSO_4 (aq). The reaction goes to completion, and a white precipitate (BaSO_4) forms. The student filters the precipitate and dries it overnight. The data are given in the following table.

Mass of dry filter paper	0.764 g
Volume of $\text{CuSO}_4(\text{aq})$	50.0 mL
Volume of 0.200 M $\text{Ba}(\text{NO}_3)_2$	20.0 mL
Mass of filter paper and dried precipitate	1.136 g

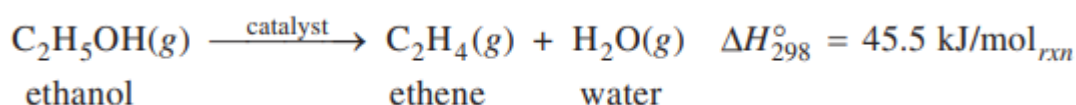
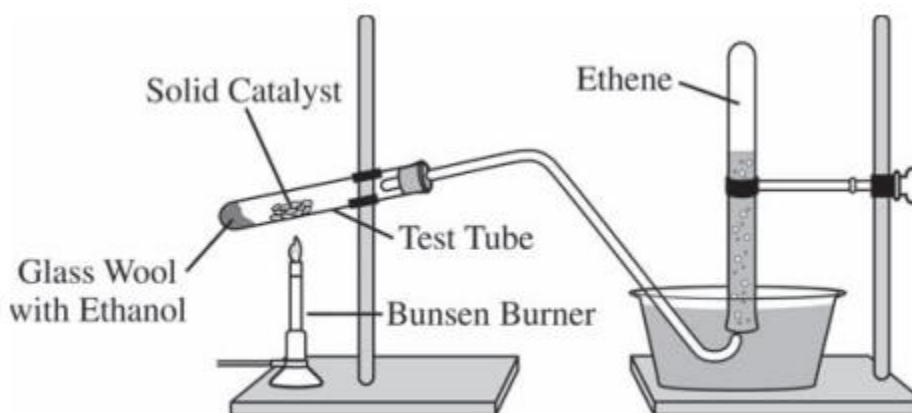
- (a) Calculate the mass of the dried precipitate, BaSO_4 (s).
- (b) Calculate the moles of the dried precipitate, BaSO_4 (s).
- (c) Determine the moles of the sulfate ion, SO_4^{2-} (aq).
- (d) Determine the moles of the original CuSO_4 (aq) solution.
- (e) Calculate the molarity of the original CuSO_4 (aq) solution in moles per Liter.



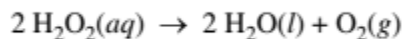
2. A student dissolved a solid sample of oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$, in water in an Erlenmeyer flask. Then the student titrated the $\text{H}_2\text{C}_2\text{O}_4$ solution in the flask with a solution of KMnO_4 , which has a dark purple color. The balanced chemical equation for the reaction that occurred during the titration is shown above. The student used a 50.0 mL buret to add the KMnO_4 (aq) to the $\text{H}_2\text{C}_2\text{O}_4$ (aq) until a faint lavender color was observed in the flask, an indication that the end point of the titration had been reached. The initial and final volume readings of the solution in the buret are shown below. Be sure to read the initial reading and final readings of the buret to two decimal places.



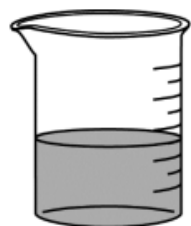
- Determine the volume of KMnO_4 (aq) in mL that was added from during the titration from your initial and final readings of the buret.
- Given that the concentration of KMnO_4 (aq) was 0.0235 M, calculate the number of moles of MnO_4^- ions that completely reacted with the $\text{H}_2\text{C}_2\text{O}_4$.
- Calculate the number of moles of $\text{H}_2\text{C}_2\text{O}_4$ (aq) that was reacted during the titration.
- Calculate the mass in grams of $\text{H}_2\text{C}_2\text{O}_4$ that was fully dissolved in the water in the Erlenmeyer flask.



3. Ethene, $\text{C}_2\text{H}_4(g)$ may be prepared by the dehydration of ethanol, $\text{C}_2\text{H}_5\text{OH}(g)$, using a solid catalyst. A setup for the lab synthesis is shown in the diagram above as well as the equation for the dehydration reaction. A student adds a solid sample of $\text{C}_2\text{H}_5\text{OH}(l)$ to a test tube and the test tube is heated gently with a Bunsen burner until all of the $\text{C}_2\text{H}_5\text{OH}(l)$ evaporated and gas generation stopped. When the reaction stopped, the volume of collected gas over the water was 0.0854 L at 0.822 atm and 305 K. (The vapor pressure of water at 305 K is 0.0470 atm)
- (a) If the gas collected over the water is gaseous water (water vapor) and gaseous ethene, $\text{C}_2\text{H}_4(g)$, calculate the pressure of just ethene, $\text{C}_2\text{H}_4(g)$, in the collection tube.
- (b) Calculate the number of moles of ethene, $\text{C}_2\text{H}_4(g)$, collected during the reaction.
- (c) Determine the number of moles of ethanol, $\text{C}_2\text{H}_5\text{OH}(g)$, which was reacted.
- (d) Calculate the amount of heat in kJ absorbed by the $\text{C}_2\text{H}_5\text{OH}(g)$ during the reaction.

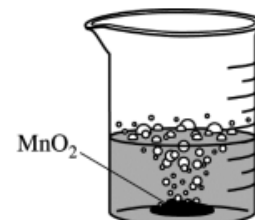


4. Hydrogen peroxide, H_2O_2 , decomposes according to the equation above. A student investigates the decomposition reaction in the laboratory. The student prepares two small beakers, adding 20.0 mL of 9.77 M H_2O_2 (aq) to each one. Each beaker is placed on an electronic balance. The student adds 0.10 g of MnO_2 (s) to the second beaker and records the mass of each beaker and its contents at 10-second intervals for one minute. The beakers and the data are shown below.



Beaker 1

Time (seconds)	Mass of Beaker 1 + $\text{H}_2\text{O}_2(\text{aq})$ (grams)	Mass of Beaker 2 + $\text{H}_2\text{O}_2(\text{aq})$ + $\text{MnO}_2(\text{s})$ (grams)
0	43.09	43.19
10	43.09	43.06
20	43.09	42.94
30	43.09	42.83
40	43.09	42.73
50	43.09	42.65
60	43.09	42.58



Beaker 2

For beaker 2 during the 60-second period:

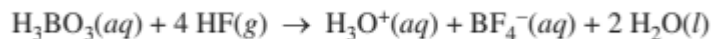
- Calculate the mass of oxygen gas, O_2 (g), produced.
- Calculate the moles of oxygen gas, O_2 (g), produced.
- Determine the number of moles of hydrogen peroxide, H_2O_2 (aq), which was decomposed.
- Calculate the number of moles of hydrogen peroxide, H_2O_2 (aq), which were initially added to the beaker.
- Calculate the percentage (using moles) of H_2O_2 (aq) that was reacted in this 60-second period.



5. $\text{NaHCO}_3(s)$ (baking soda) decomposes upon heating to produce $\text{Na}_2\text{CO}_3(s)$ and two gaseous products, as shown by the equation above. A student conducts an experiment to determine the composition of a mixture of NaHCO_3 and Na_2CO_3 . The student places a sample of the mixture into a preweighed test tube that is attached to a container that holds a drying agent. The student heats the test tube strongly with a Bunsen burner for 10 minutes, during which time all of the water produced by the reaction is captured by the drying agent. The following table shows the data the student recorded during the experiment.

Mass of empty test tube	15.825 g
Mass of test tube and mixture before heating	17.648 g
Mass of drying agent before reaction	2.134 g
Mass of drying agent and water after reaction	2.303 g

- (a) Calculate the mass of water collected by the drying agent during the reaction.
- (b) Calculate the number of moles of water collected by the drying agent during the reaction.
- (c) Determine the number of moles of $\text{NaHCO}_3(s)$ present in the mixture in the test tube before the reaction was initiated.
- (d) Calculate the number of grams of $\text{NaHCO}_3(s)$ present in the mixture in the test tube before the reaction was initiated.
- (e) Calculate the number of grams of the mixture before heating.
- (f) Calculate the mass percent of $\text{NaHCO}_3(s)$ in the mixture.



6. Tetrafluoroboric acid is a strong acid with the formula HBF_4 . The acid can be prepared by reacting the weak acid H_3BO_3 with HF according to the equation above. To prepare a solution of this strong acid, $\text{HF}(g)$ is bubbled into a solution containing 50.0 g of H_3BO_3 in a 200 mL reaction vessel.
- (a) Calculate the maximum number of moles of H_3BO_3 that can be reacted in the vessel.
- (b) Determine the maximum number of moles of $\text{HF}(g)$ that can be used during the reaction.
- (c) Calculate the number of liters of $\text{HF}(g)$, measured at 273 K and 1.00 atm, that will be consumed if all the H_3BO_3 reacts.
- (d) Determine the maximum number of moles of $\text{H}_3\text{O}^+(aq)$ that can be produced.
- (e) If the total volume of the products are measured to be 31.2 mL, calculate the concentration of the strong acid, $\text{H}_3\text{O}^+(aq)$, in the reaction vessel.