

How to Prepare a Solution

Many experiments involving chemicals call for their use in solution form. That is, two or more substances are mixed together in known quantities. This may involve weighing a precise amount of dry material or measuring a precise amount of liquid. Preparing solutions accurately will improve an experiment's safety and chances for success.

You prepare a solution by dissolving a known mass of solute (often a solid) into a specific amount of a solvent. One of the most common ways to express the concentration of the solution is **M** or molarity, which is a **mole** of solute per liter of **solution**.

Preparing Chemical Solutions

Lab experiments and types of research often require preparation of chemical solutions in their procedure. We look at preparation of these chemical solutions by weight (w/v) and by volume (v/v). The glossary below cites definitions to know when your work calls for making these and the most accurate molar solutions.

To this we add information designed for understanding how to use the **pH scale** when measuring acidity or alkalinity of a solution.

Glossary, basic terms to understand...

Solute - The substance which dissolves in a solution

Solvent - The substance which dissolves another to form a solution. For example, in a sugar and water solution, water is the solvent; sugar is the solute.

Solution - A mixture of two or more pure substances. In a solution one pure substance is dissolved in another pure substance homogeneously. For example, in a sugar and water solution, the solution has the same concentration throughout, ie. It is homogeneous.

Mole - A fundamental unit of mass used by chemists. This term refers to a large number of elementary particles (atoms, molecules, ions, electrons, etc...) of any substance.

Example of How to Prepare a Solution

Prepare **1 liter** of **1.00 M** NaCl solution.

First, calculate the **molar mass** of NaCl which is the mass of a mole of Na plus the mass of a mole of Cl or $22.99 + 35.45 = 58.44 \text{ g/mol}$.

1. Weigh out 58.44 g NaCl.
2. Place the NaCl in a **1-liter** volumetric flask.
3. Add a small volume of distilled, deionized water to dissolve the salt.
4. Fill the flask to the **1 L** line.

If a different **molarity** is required, then multiply that number times the molar mass of NaCl. For example, if you wanted a **0.5 M** solution, you would use $0.5 \times 58.44 \text{ g/mol}$ of NaCl in **1 L** of solution or 29.22 g of NaCl.

Important Points to Remember

- **Molarity** is expressed in terms of **liter of solution**, not liters of solvent. To prepare a solution, the flask is filled to the mark. In other words, it is incorrect to add 1 liter of water to a mass of sample to prepare a molar solution.
- Sometimes it's necessary to adjust **the pH of a solution**. To do this, add enough water to dissolve the solute. Then add an acid or base solution drop wise (usually a hydrochloric acid or HCl solution for acid or sodium hydroxide or NaOH solution for a base) to reach the desired pH. Then add more water to reach the mark on the glassware. Adding more water won't change the pH value.

Solution 1: Using percentage by weight (w/v)

Formula

The formula for weight percent (w/v) is: $[\text{Mass of solute (g)} / \text{Volume of solution (ml)}] \times 100$

Example

A **10% NaCl** solution has ten grams of sodium chloride dissolved in 100 ml of solution.

Procedure

Weigh **10g** of sodium chloride. Pour it into a graduated cylinder or volumetric flask containing about **80ml** of water. Once the sodium chloride has dissolved completely (swirls the flask gently if necessary), add water to bring the volume up to the final 100 ml. Caution: Do not simply measure **100ml** of water and add 10g of sodium chloride. This will introduce error because adding the solid will change the final volume of the solution and throw off the final percentage.

Solution 2: Using percentage by volume (v/v)

When the solute is a liquid, it is sometimes convenient to express the solution concentration as a volume percent.

Formula

The formula for volume percent (v/v) is: $[\text{Volume of solute (ml)} / \text{Volume of solution (ml)}] \times 100$

Example

Make **1000ml** of a 5% by volume solution of ethylene glycol in water.

Procedure

First, express the percent of solute as a decimal: $5\% = 0.05$

Multiply this decimal by the total volume: $0.05 \times 1000\text{ml} = 50\text{ml}$ (ethylene glycol needed).

Subtract the volume of solute (ethylene glycol) from the total solution volume:

$1000\text{ml (total solution volume)} - 50\text{ml (ethylene glycol volume)} = 950\text{ml}$ (water needed)

Dissolve **50ml** ethylene glycol in a little less than **950ml** of water. Now bring final volume of solution up to **1000ml** with the addition of more water. (This eliminates any error because the final volume of the solution may not equal the calculated sum of the individual components).

So, 50ml ethylene glycol / 1000ml solution x100 = 5% (v/v) ethylene glycol solution.

Solution 3: Molar Solutions

Molar solutions are the most useful in chemical reaction calculations because they directly relate the moles of solute to the volume of solution.

Formula

The formula for molarity (**M**) is: moles of solute / 1 liter of solution or gram-molecular masses of solute / 1 liter of solution.

Examples

The molecular weight of a sodium chloride molecule (NaCl) is 58.44, so one gram-molecular mass (=1 mole) is 58.44 g. We know this by looking at the **periodic table**. The atomic mass (or weight) of Na is 22.99, the atomic mass of Cl is 35.45, so $22.99 + 35.45 = 58.44$.

If you dissolve **58.44g** of NaCl in a final volume of **1 liter**, you have made a **1M NaCl** solution, a **1 molar** solution.

Procedure

To make molar NaCl solutions of other concentrations dilute the mass of salt to **1000ml** of solution as follows:

0.1M NaCl solution requires **0.1 x 58.44 g of NaCl = 5.844g**

0.5M NaCl solution requires **0.5 x 58.44 g of NaCl = 29.22g**

2M NaCl solution requires **2.0 x 58.44 g of NaCl = 116.88g**