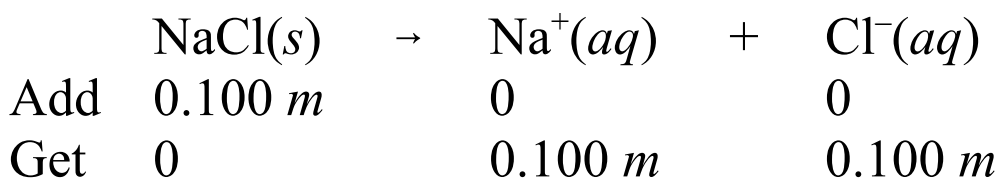


## Ideal Electrolyte Solutions

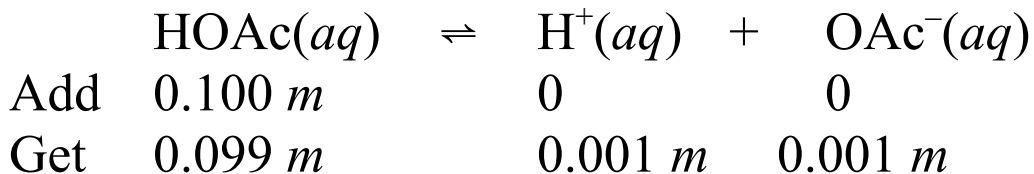
- ✓ If interactions between solute particles are minimal, the effective molality or molarity of an electrolyte solution will be the sum of the concentrations of particles from all sources; i.e., ions and any undissociated molecules (in the case of weak electrolytes).

Strong electrolyte: 0.100 *m* NaCl



$$\Rightarrow m_t = 0.200 \text{ } m$$

Weak electrolyte: 0.100 *m* HOAc



$$\Rightarrow m_t = 0.101 \text{ } m$$

## van't Hoff $i$ Factor

- ☞ In real solutions, ions interact (non-ideal behavior), making their *effective concentrations* and the associated colligative effects less than expected.
- ☞ An empirical measure of the dissociation of a solute is the *van't Hoff factor*,  $i$ , defined by any of the following:  
$$\Delta T_f = imK_f \quad \Delta T_b = imK_b \quad \pi = iMRT$$
- ☞ If ions did not interact in electrolyte solutions (ideal behavior), for *strong electrolytes* ideal values of  $i$  would be equal to the number of ions per formula unit.

Electrolyte	Ideal $i$
NaCl	2
K <sub>2</sub> SO <sub>4</sub>	3
(NH <sub>4</sub> ) <sub>3</sub> PO <sub>4</sub>	4

## van't Hoff $i$ Values of Real Solutions

- Real solutions of strong electrolytes have observed values of  $i$  that are less than the ideal values, due to inter-ionic interactions.

$$i_{\text{observed}} < i_{\text{ideal}}$$

- Observed  $i$  values become less ideal with greater concentration and more ideal with greater dilution.

Concentration $\text{K}_2\text{SO}_4$	Value of $i$
0.100 $m$	2.32
0.0100 $m$	2.69
0.00100 $m$	2.84

- Weak electrolytes, which produce relatively few ions in solution, have observed  $i$  values slightly greater than 1, the ideal value for a non-electrolyte.