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

	$NaCl(s)$	→	$Na^+(aq)$	+	$Cl^-(aq)$
Add	0.100 *m*		0		0
Get	0		0.100 *m*		0.100 *m*

$$\Rightarrow m_t = 0.200\ m$$

Weak electrolyte: 0.100 *m* HOAc

	$HOAc(aq)$	⇌	$H^+(aq)$	+	$OAc^-(aq)$
Add	0.100 *m*		0		0
Get	0.099 *m*		0.001 *m*		0.001 *m*

$$\Rightarrow m_t = 0.101\ m$$

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:

$$\Delta T_f = imK_f \qquad \Delta T_b = imK_b \qquad \pi = iMRT$$

☞ 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_2SO_4	3
$(NH_4)_3PO_4$	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 K_2SO_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.