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Ag (I) Cationization Mass Spectrometry:
Substituent Effects in Cation-pi Interactions

A Thesis Submitted in Partial Fulfillment of the
Requirements of the Degree of

MASTER OF PHILOSOPHY

in

Chemistry

by

Wong Yuen Pik

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Abstract

Cation-$\pi$ interaction has been recently recognized as a new type of non-covalent interaction important in biological recognition, and the design of functional materials in the nanoscale or sub-nanoscale. The binding Ag$^+$ ion to alkylbenzenes (RBz), substituted naphthalenes (X-Nap, where $X = H$, Me, Et, i-Pr, OH, OMe, OEt, CN, NH$_2$), substituted indoles (X-Indole, where $X = H$, Me, OH, OMe, NO$_2$, CN) and substituted phenols (X-PhOH, where $X = H$, Me, Et, i-Pr, t-Bu, OMe, OEt, NO$_2$) are chosen as model cation aromatic-$\pi$ systems to study substituent effects on cation-$\pi$ binding, as well as the factors governing the strength of cation-$\pi$ interactions.

The Ag$^+$ binding affinities of these aromatic model ligands (L) at 0K were measured by vigorous application of the mass spectrometric kinetic method, which is based on the competitive dissociation of the Ag$^+$ bound heterodimers complexes, [L$_1$ + Ag + L$_2$]$^+$, to their respective monomer complexes. The Ag$^+$ bound heterodimers were generated by electrospray ionization (ESI), and their dissociations were conducted under low-energy (Ar as collision gas) and / or high-energy (He as collision gas) collision-induced dissociation (CID) conditions.

The Ag$^+$ binding affinities of alkylbenzenes (157 – 216 kJmol$^{-1}$), substituted naphthalene (177 – 211 kJmol$^{-1}$), substituted indoles (201 – 225 kJmol$^{-1}$), and substituted phenols (158 – 185 kJmol$^{-1}$) were found to be in good agreement (experimental uncertainty of ± 10 – 14 kJ mol$^{-1}$) with theoretical $ab$ initio affinities estimated at the CCSD(T)/[HW(f), 6-31G+(d)] level.

Our experimental and theoretical studies show that Ag$^+$ cation-$\pi$ binding is the most stable and preferred binding mode for alkyl, hydroxyl and alkoxy substituted...
benzenes, naphthalenes, indoles and phenols, indicating that non-covalent cation-π interaction is indeed strong enough to compete against Ag⁺ binding to oxygen-donor binding sites. On the other hand, Ag⁺ cation-π binding is not as strong as Ag⁺ binding to heteroatom nitrogen / oxygen binding sites in –CN, –NH₂ and –NO₂ substituted benzenes, naphthalenes, indoles and phenols. Our results indicate that Ag⁺-π interaction is mainly electrostatic in nature, though charge-transfer covalent interaction is noticeably present. The relative stability of the cation-π and non-π binding to O/N heteroatom sites is mainly determined by the interplay of electrostatic ion-quadrupole, ion-dipole, ion-induced dipole interactions, as well as the extent of charge-transfer (covalent) interaction present in the Ag⁺-ligand complexes.

Ag⁺ binding to substituted naphthalenes and substituted indoles are measured to be ~ 9 – 12 % and ~ 11 – 33 % stronger than of substituted benzenes, respectively, indicating that both the quadrupole moment and polarizability of fused aromatic ring systems could greatly enhance cation-π binding energy. Substituted phenols show comparable Ag⁺ affinities as substituted benzenes, i.e., the phenolic-OH group does not have any significant effect on Ag⁺ cation-π binding affinities. The increasing trend of Ag⁺ affinities within a series of substituted benzenes, naphthalenes, indoles and phenols is shown to be mainly due to the increasing molecular polarizability of the substituted aromatic systems, which could enhance the Ag⁺ affinities via both ion-induced dipole and charge-transfer interactions.

Ag⁺ cation-π binding affinities are comparable to that of Li⁺, but greater than Na⁺ affinities, even though the ionic radius of Ag⁺ is larger than that of Na⁺. The greater Ag⁺ affinity is attributed to the noticeable presence of charge-transfer
(covalent) interactions, i.e., $\pi \rightarrow \text{Ag}^+$ ($\sigma$-donor interaction) and $\text{Ag}^+ \rightarrow \pi^*$ ($4d^{10}$ back donation), present in $\text{Ag}^+$-ligand interactions. In fact, the order of transitional metal cation-$\pi$ affinities: $\text{Co}^+ > \text{Fe}^+ > \text{Cr}^+ \sim \text{Ag}^+$ is rationalized in terms of the variation in charge-transfer interaction for different transition metal cations.
Certificate of Originality

I hereby declare that the work described in this thesis were carried out by me during the period Sept 2001 to Oct 2003 in the Department of Applied Biology and Chemical Technology, The Hong Kong Polytechnic University. All *ab initio* M.O. calculations presented in this thesis were carried out by Miss H. M. Lee of our research group. This thesis has not been submitted to this or other institution for any academic qualification.

________________________
Wong Yuen Pik

Date: 24th October, 2003
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List of Abbreviations

ACN  acetonitrile
B    magnetic sector
BSSE basic-set-superposition-error
C.L. confidence interval
CID  collision-induced dissociation
CTST canonical transition state theory
DFT  density functional theory
E    electric sector
ESA  electrostatic analyzer
ESI  electrospray ionization
FFR  field free region
HPMS High-pressure mass spectrometry
ICR  ion cyclotron resonance
m/z  mass-to-charge ratio
MAD  mean absolute deviation
MIKE mass-analyzed ion kinetic energy
M.O. molecular orbital
MS   mass spectrometry
MS/MS tandem mass spectrometry
M.W. molecular weight
NBO  natural bond orbital
NPA  natural population analysis
PAH  polynuclear aromatic hydrocarbon
RBz  alkylbenzenes
QET  quasi-equilibrium theory
SD   standard deviation
T_{eff} effective temperature
TOF  time-of-flight
X-Indole substituted indoles
X-Nap substituted naphthalenes
X-PhOH substituted phenols
ZPVE zero point vibrational energy
Chapter 1  Introduction

1.1  Biochemical Significance of Cation-π Interaction

Non-covalent interactions are important in many areas of modern chemistry, such as the design of new materials and molecular recognition in synthetic host-guest receptor complexes. Cation-π interaction is the non-covalent binding interaction between a cation and the π face of an aromatic structure.

In the past years, there are increasing evidence suggesting that cation-π interactions are important in determining the structure-functional aspects of proteins in biological systems [Ma and Dougherty, 1997; Gokel et al., 2000; Zacharias and Dougherty, 2002; Kim et al., 2000; 2003]. Recent examples include the implications of cation-π interaction in the biological function of acetylcholine esterase [Harel et al., 1993], trimethylamine dehydrogenase [Basran et al., 1997], liver glycogen phosphorylase inhibitor [Rath et al., 2000], voltage-gated Na⁺ channels for local anaesthetic [Wright et al., 1998; Li et al., 1999], MP2 ion channel of influenza A virus [Okada et al., 2001], and the enhanced stability of cell membrane proteins [Yau et al., 1998].

Cation-π interaction has been exploited in the design of therapeutical agents against various ailments like myasthenia gravis, glaucoma and Alzheimer’s disease [Krogsgard-Larsen, 1991], and in the design of collarenes for separating $^{137}$Cr and $^{90}$Sr from nuclear wastes [Choi et al., 1998]. A detailed understanding of cation-π interaction would aid in the design of synthetic host-guest receptors [De Wall, 2000; Oh et al., 2000] and novel functional materials, such as metallic nanostructures [Hong et al., 2001].
A key parameter determining the occurrence and properties of cation-π interaction is the cation binding energy / affinity (enthalpy of binding, ΔH) between the cation and the aromatic ligand. The cation (M⁺) binding affinity of an aromatic ligand molecule (L) is defined as the enthalpy change of Reaction 1.1, and corresponds to the binding energy of the M⁺-L complex at 0 K, which is calculated by Equation 1.2.

\[
M⁺-L \rightarrow M⁺ + L \quad [1.1]
\]

\[
ΔH = E_{M⁺} + E_L - E_{M⁺-L} \quad [1.2]
\]

where M⁺ is the cation (e.g. Ag⁺ in the present study), L is an aromatic organic ligand, and \(E_{M⁺}\), \(E_L\) and \(E_{M⁺-L}\) are the total energies of M⁺, L and M⁺-L. The binding affinities at 0 and 298K are often denoted as ΔH₀ and ΔH₂₉₈, respectively.

The strength of the cation-π interaction is generally greater than that of other non-covalent binding forces (e.g. π-π, π-hydrogen bonds), but may vary significantly, depending on the nature of the cation and the aromatic-π surface. For metal cations, the cation-π binding energies are estimated in the range of 100 – 300 kJ mol⁻¹ (1/7 to 1/2 of the strength of a C-C bond) [Ma and Dougherty, 1997; Kim et al., 2000; 2003], which are not easily measured by conventional spectroscopic techniques. Furthermore, cation-π interaction cannot be easily isolated from other non-covalent interactions present in the more complex, multi-functional chemical and biochemical systems. Up to now, most theoretical and experimental studies on cation-π interaction have been conducted with aromatic model compounds [Caldwell and Kollman, 1995; Kim et al., 1994; Dougherty, 1996; Zhu et al., 2000; Cubero et al., 1999; Mecozzi et al., 1996; Quínonero et al., 2002] or synthetic host-guest receptors.
[Gokel, 2000]. A similar approach of using model substituted aromatic ligands is adopted in the present study as well.

1.2 Mass Spectrometric Studies on Metal Cation-π Interactions in the Gas Phase

Studies of cation-ligand interactions in the gas phase have the distinct advantage of elucidating the *intrinsic* nature of the binding interaction, in the absence of complicating effects of the solvent. Such studies are usually performed in the vacuum environment of a mass spectrometer. Given the importance of alkali metal cations in biological systems, the binding affinities of alkali metal cations with benzene and a limited number of substituted derivatives have been determined recently by the threshold collision-induced dissociation technique [Amicangelo and Armentrout, 2000; Ammunugama and Rodgers, 2002a; 2002b; 2002c; 2003b; 2003c]. However, the effects of substituents on cation-π binding affinities have not been fully investigated.

Although not as biochemical relevant as alkali metal cations, Ag⁺ ion has long been found to bind strongly to biomolecules such as amino acids [Lee et al., 1998; Talaty et al., 2001; Shoeib et al., 2001a] and peptides [Li et al., 1997; Chu et al., 2001; Shoeib et al., 2001b]. The Ag⁺ ion is used as a bactericide in eye drops for newborns [Petering, 1976; Wigley and Brooks, 1992], and some silver complexes have shown remarkable antimicrobial activities [Nomiya et al., 1995a; 1995b]. Much of the biological action is believed to be initiated by Ag⁺ binding to functional group of proteins in microorganisms [Petering, 1976; Fox and Modak, 1974]. For example, a
recent study of cystein-mutated potassium ion channel in Murine Kir 2.1 showed that the H5 pore region of the channel could be blocked by Ag⁺ [Dart et al., 1998]. Furthermore, mass spectral fragmentation of Ag⁺-peptide complexes has been demonstrated to be useful in peptide sequencing [Chu et al., 1999; 2002]. The competition among different fragmentation pathways are determined, to a large extent, by the energetics of Ag⁺ binding to the different binding sites of the peptide, including Ag⁺ cation-π binding to aromatic phenyl ring of phenylalanine.

In this study, we chose the Ag⁺ ion as the 'probe' cation to investigate the effect of different types of substituents in Ag⁺ cation-π interaction. Ag⁺ is chosen because of feasibility on both experimental and theoretical studies. Compared to other transitional metal ions, Ag⁺ is relatively inert in the gas phase. This is due to its stable [Kr](4d¹⁰) electronic configuration [Jones and Staley, 1982; Weil and Wilkins, 1985], and the large difference in energy between its ground and first excited state ([Kr] (4d⁸)(5s¹) at 4.9 eV) [Jones and Staley, 1982], or the Ag²⁺ state (second ionization potential of Ag at 21.5 eV) [Walker et al., 1999], which can affect its binding nature / energies towards different ligands. Previous results in our laboratory showed that stable Ag⁺ bound heterodimer of benzene and related aromatic π-systems could be easily and abundantly generated by electrospray ionization techniques [Ma et al., 1997; Ng et al., 1998; Ng, 1999]. This easy way of preparing Ag⁺ bound heterodimers of aromatic ligands allows us to determine their Ag⁺ affinities by the kinetic method (refer to Section 2.4) with conventional tandem mass spectrometric systems. Studies with other and more biological relevant transitional metals, such as Cu²⁺/Cu⁺ and Fe²⁺/Fe⁺, are more difficult because they are more reactive in the gas phase.
Studies on non-covalent interactions are often augmented by molecular orbital calculations, which provide complimentary information on the binding geometries / modes and rationalization of experimental observations. In recent years, advances are made in theoretical studies on the binding geometries and nature of binding of many different Ag⁺-ligand complexes [Bauschlicher et al., 1989; 1992; Ma et al., 1997; Ma, 1998; Boutreau et al., 2001; Kim et al., 2003]. These theoretical studies are useful in different contexts: (i) to complement and validate experimental binding affinity results, (ii) to address the experimentally difficult question of the relative stabilities of different metal cation binding sites, and (iii) to provide molecular properties (e.g. molecular quadrupole moment, polarizability, dipole moment) for the understanding of cation-π interaction. In the present study, these complimentary calculations were carried out by Miss H.M. Lee of our research group.

The binding of cations to aromatic-π face of benzene, phenol and indole, which are model systems that resembles the side chains of the amino acids phenylalanine (Phe), tyrosine (Tyr) and tryptophan (Trp), respectively, are of fundamental interest for investigating the cation-π interaction in biological systems. Also, the strength and properties of the cation-π interaction can be manipulated by substituting functional groups (substituents) to the aromatic ring. The present study focuses principally on the experimental measurement of the binding energies (affinities) between the Ag⁺ ion and alkylbenzenes / substituted naphthalenes (fused benzene systems analogous to indole) / indoles / phenols using the mass spectrometric kinetic method. For substituted aromatic compounds, another related question is the effect of substituents on the competition between cation-π and alternative non-π binding sites for substituents containing an oxygen / nitrogen heteroatom co-ordination site. In other words, is cation binding to the π-face of the aromatic ring or to the O/N
heteroatom of the functional group / aromatic substituent more stable and energetically preferred?

1.3 Non-Covalent Electrostatic Interactions

Dewar first suggested in 1951 that the strong binding of Ag⁺ to double bond(s) was due to the two-way, donor-acceptor (charge-transfer) interactions illustrated in Figure 1.1 [Herberhold, 1974]. For Ag⁺-olefin complexes, it is suggested that the Ag⁺ ion is located in the plane of the π-electrons, and on top of the C=C double bond. The major bonding interactions are covalently based: (i) "σ-type bond" arising from the transfer of π-electrons from the occupied 2p bonding orbitals of olefin into the vacant 5s orbital of the Ag⁺ ion (Figure 1.1a), and (ii) "π-type bond" arising from back-donation of 4d electrons of the Ag⁺ ion into the unoccupied π* 2p anti-bonding orbitals of the ligand (Figure 1.1b).

![Diagram of bonding interactions](image)

(a) σ-type bond  
(b) π-type bond

Figure 1.1 Two-way donor-acceptor interactions in Ag⁺-olefin systems.

However, recent theoretical studies on Ag⁺-ligand complexes in the gas phase suggest that Ag⁺ binding is mainly electrostatic in nature, although charge-transfer (covalent) may be present [Baushlicher et al. 1992; Ma, 1997; Dargel et al., 1999;
Boutreau et al., 2001; Kim et al, 2003]. The electrostatic interaction can be derived from ion-quadrupole, ion-induced dipole and ion-dipole interactions.

**Ion-Quadrupole Interaction.** Molecular quadrupole moment is a measure of the non-uniform distribution of charge within a molecule, relative to a particular molecular axis [Williams, 1993]. Benzene is a prototypical aromatic-π system that has no dipole moment but a permanent, non-spherical charge distribution or quadrupole moment that is expected to interact with appropriately positive charges through electrostatic interactions (refer to Footnote 4.1). Dougherty and co-workers first suggested that ion-quadrupole interaction and cation-π interaction is of major importance in molecular recognition in synthetic or biological receptors [Dougherty, 1996; Ma and Dougherty, 1997].

It is important to note that there is no priori reason to expect the cation-π interaction should be inherently weaker than that of ion-dipole interaction. Classically, ion-quadrupole interaction is expected to show a \(1/r^3\) dependence of stabilization energy [Williams, 1993; Dougherty, 1996] and the intermolecular potential, \(\tau_{\text{ion-quadrupole}}\), of ion-quadrupole is shown as Equation 1.3 by [Israchelvili, 1992].

\[
\tau_{\text{ion-quadrupole}} = \frac{Qq(3\cos^2\Phi - 1)}{4r^3} \tag{1.3}
\]

where \(\tau_{\text{ion-quadrupole}} = \text{ion-quadrupole interaction potential (in Volts, V),}\)

\(Q = \text{electric charge of the ion (in Coulomb, C),}\)

\(q = \text{point charge of the ligand (in Coulomb, C),}\)

\(\Phi = \text{angle of deviation between the ion from the quadrupole moment vector (in degrees),}\)
\( r = \) distance between ion and the centre of the quadrupole axis (in meter, m).

In practice, cation-\( \pi \) interaction rather exhibits \( 1/r^n \) (with \( n < 2 \)) dependence [Ma and Dougherty, 1997] because other types of interactions such as ion-induced dipole and charge-transfer interactions are also present [Ma, 1997; Boutreau et al., 2001; Kim et al., 2003].

**Ion-Induced Dipole Interaction.** Ion induced-dipole interaction has been suggested to be one of the major contributions to electrostatic cation-\( \pi \) interaction [Ma, 1997; Ng et al., 1998]. Classically, ion-induced dipole interaction would occur when a charge is separated at a distance from the molecular centre of polarizabilities. The interaction energy is given by the following equation [Isracelacvili, 1992]:

\[
E_{\text{ion-induced dipole}} = -\frac{Q^2\alpha}{2(4\pi\varepsilon_0)^2 r^4} \tag{1.4}
\]

where \( E_{\text{ion-induced dipole}} = \) ion-induced dipole interaction energy (in Joules, J),

- \( Q = \) electric charge of the ion (in Coulombs, C),
- \( \alpha = \) polarizability of the ligand (in Coulombs-meter, Cm),
- \( r = \) distance between the interacting ion and ligand (in meter, m),
- \( \varepsilon_0 = \) dielectric permittivity of free space \( (8.854 \times 10^{-12} \text{ C}^2\text{J}^{-1}\text{m}^{-1}) \).

A metal cation approaching a polarizable molecule can induce an electric dipole moment in a neighboring ligand, no matter the molecule is polar or not. As shown in Equation 1.4, the ion-induced dipole interaction energy depends on the polarizability of the ligand, and is inversely proportional to \( r^4 \).

**Ion-Dipole Interaction.** Ion-dipole interaction is the attractive electrostatic interaction between the electric charge of the ion and the dipole moment, \( \mu \), of the
ligand. The electrostatic interaction energy attributed to ion-dipole interaction, $E_{\text{ion-dipole}}$, is classically given by Equation 1.5 [Isracelachvili, 1992].

$$E_{\text{ion-dipole}} = -\frac{Q \mu \cos \Phi}{4\pi \varepsilon_0 r^2}$$ [1.5]

where $E_{\text{ion-dipole}} =$ ion-dipole interaction (binding) energy (in Joules, J),

$Q =$ electric charge of the ion (in Coulomb, C),

$\mu =$ dipole moment of the ligand (in Coulomb-meter, Cm),

$\Phi =$ angle of deviation between the ion from the dipole moment vector (in degrees),

$r =$ distance between ion and the centre of the dipole axis (in meter, m),

$\varepsilon_0 =$ dielectric permittivity of free space ($8.854 \times 10^{-12} \text{ C}^2 \text{J}^{-1} \text{m}^{-1}$).

The classical descriptions of electrostatic interactions are useful in a conceptual and qualitative way to explain the observed Ag$^+$ binding affinities. However, the equations of classical electrostatics fail to predict or estimate quantitatively the overall electrostatic binding energies. At present, the relative contributions of ion-quadrupole, ion-induced dipole, ion-dipole and charge-transfer interactions to Ag$^+$-ligand binding, and in particular, to cation-π interaction, is not well understand. However, the common consensus is that covalent (charge-transfer) interaction, though minor, is still important in affecting the binding geometry and to a lesser extent, the affinities of Ag$^+$-ligand complexes.
1.4 Scope and Objectives of the Present Study

The aim of the present study is to determine the gas phase Ag⁺ affinities of selected classes of substituted aromatic ligands (Figure 1.2) by the mass spectrometric kinetic method. High-level ab initio molecular orbital calculations were also conducted to aid in the interpretation of the experimental results and for providing complimentary information on the physico-chemical factor(s) governing the binding energies / geometries / nature of Ag⁺ cation-π and non-π modes of binding. The results of the present study are also of useful reference to the binding of other biologically relevant transitional metals, such as Cu²⁺/Cu⁺ and Fe²⁺/Fe⁺, to aromatic amino acid residues in the gas phase.

To summarize, the specific objectives of our present study are:

1. To determine the experimental Ag⁺ binding affinities of different aromatic model systems such as alkylbenzenes (RBz, Chapter 4), substituted naphthalenes (X-Nap, Chapter 5), substituted indoles (X-Indole, Chapter 6) and substituted phenols (X-PhOH, Chapter 7) by the mass spectrometric kinetic method;

2. To investigate the effects of different substituents on the binding strength of cation-π interactions, and the relative importance of electrostatic and charge-transfer interactions in Ag⁺ binding to these model aromatic ligands; and

3. To compare the binding affinities of Ag⁺-aromatic model compounds with that of alkali metal cations (Li⁺, Na⁺, K⁺, Rb⁺ and Cs⁺) and other transition metal cations (Cu⁺, Au⁺, Cr⁺, Fe⁺ and Co⁺) so as to elucidate the cationic effects on cation-π interactions.
Figure 1.2  Chemical structures of substituted aromatic model systems used in the present study.
Chapter 2  Mass Spectrometric Techniques and Related Topics

2.1  Electrospray Ionization (ESI)

In our laboratory, we have found that abundant Ag$^+$ bound monomer and heterodimer complexes with aromatic ligands such as benzene, alkylbenzenes and polyaromatic hydrocarbons could be easily generated by eletrospray ionization (ESI). ESI, introduced by Yamashita and Fenn [Yamashita and Fenn, 1984], is a soft ionization technique that allows the transfer of ions from solution to the gas phase, from which the ions are then subjected to mass spectrometric analysis. Neutral compounds can also be converted to the ionic form in solution and/or in the gas phase via ion-molecule reactions or metal cationization.

A schematic diagram of an ESI interface and the ion production processes is shown in Figure 2.1. There are three major steps in the production of gas phase ions from sample solutions by ESI: (1) production of charged droplets at the electrospray capillary tip applied with high voltage; (2) shrinkage of the charge droplets by solvent evaporation and repeated droplet disintegrations, leading ultimately to very small highly charged droplets capable of producing gas-phase ions; and (3) the actual mechanism by which gas-phase ions are produced from the very small and highly charged droplets. The detailed mechanisms involved in the electrospray process is quite complicated, and have been reviewed by Kebarle and co-workers [Kebarle and Tang, 1993; Kebarle and Ho, 1997; Kebarle, 2000].
ESI source is now commonly available on magnetic sector, quadrupole, ion trap and time-of-flight (TOF) mass spectrometers. Compared to other ionization methods of generating inorganic and organometallic complexes for mass spectrometric analysis, such as glow-discharge, laser vaporization, thermionic emission and ion beam sputtering, ESI is clean, simple and applicable to a wide range of polar, thermolabile and ionic samples.
2.2 The Quasi-Equilibrium Theory (QET) of Mass Spectrum

In the simplest term, a mass spectrum is a record of the relative ion abundances of a precursor ion and its fragment ions resulting from unimolecular dissociation of the precursor ions. A mass spectrometer is usually operated under high vacuum such that no reactive collisions would occur between the precursor ion and a reactant molecule, i.e. the internal energy of gas phase molecular ions may not follow a Maxwell-Boltzmann distribution. The classical first order rate law applies here:

\[- \frac{d[m_1]}{dt} = \frac{d[m_2]}{dt} = -k[m_1] \quad [2.1]\]

where \(k\) is the rate constant, \(t\) is the reaction time, \(m_1\) is the parent ion, \(m_2\) is the product ion, and \([m_1]\) and \([m_2]\) denote their real-time concentrations, respectively. The time frame for magnetic sector and quadrupole mass spectrometers is very short (~10^{-5} s), and the fragment ion intensity is governed by the kinetics of the fragmentation reactions.

The reaction pathway is postulated to pass through a transition state, which is the rate-determining step of the unimolecular dissociation, as shown in Figure 2.2. The rate constant, \(k\), can be described by the canonical transition state theory (CTST):

\[ k = \frac{RT}{h} \frac{Q^*}{Q} e^{-\epsilon_0/RT} \quad [2.2]\]

where \(\epsilon_0\) is the critical (threshold) energy of the reaction, \(h\) is the Planck's constant, \(R\) is the gas constant, \(T\) is the temperature, \(Q^*\) and \(Q\) are the partition functions for the transition state and the precursor ion, respectively. The partition function is the number of microstates accessible by the ions; therefore, the ratio \(Q^*/Q\) measures the
entropy change of the transition state relative to the precursor ion. This equation is valid only when the system is under thermodynamic equilibrium, such that the population of the transition state can be described by a Boltzmann distribution with a temperature $T$. Since the ions are isolated under the vacuum conditions of a mass spectrometer, energy cannot be exchanged freely between ions; i.e., the ions are not decomposing under thermodynamic equilibrium conditions. An useful approximation to side-step this problem is to replace the thermodynamic equilibrium temperature $T$ with an 'effective temperature' term, $T_{\text{eff}}$, which is defined to be the temperature at which the population of dissociating ions would give the identical fragmentation behavior observed.

![Potential Energy Diagram](image)

**Figure 2.2** Potential energy diagram of a fragmentation reaction with a critical energy ($\varepsilon_0$).

Another method is to formulate the rate constant $k$ as a function of the total internal excitation energy, $E$. According to the quasi-equilibrium theory (QET) of mass
spectra [Levensen, 1978], the rate constant, \( k(E) \), of an unimolecular reaction can be expressed (in the simplest form) as:

\[
k(E) = \nu \left( \frac{E - e_0}{E} \right)^{s-1}
\]

[2.3]

where \( k(E) \) denotes the rate constant at an energy \( E \), \( \nu \) is the frequency factor, which is the ratio of the vibrational frequency of the reactant to that of the activated complex, and \( s \) is the 'effective number of oscillators' in the molecules (or degree of freedom), which is usually taken to be 1/2 to 1/5 of the actual vibrational degrees of freedom \((3N - 6)\) of the precursor ion. Thus, if the internal energy, \( E \), of the selected precursor ion is lower than the critical energy of a dissociation reaction (i.e., \( E < e_0 \)), then the precursor ion will not be able to undergo dissociation.

The QET of mass spectra is useful in providing a semi-quantitative model to understand the many physico-chemical parameters that affect the rate constant, \( k \), of a fragmentation reaction. In conjunction with the internal energy distribution, \( P(E) \), of the parent ions, it can determine theoretically the relative ion abundance observed in a mass spectrum. The quasi-equilibrium theory of mass spectra and the effect of \( P(E) \) on relative ion abundance has been reviewed recently [Vekey, 1996].
2.3 Tandem Mass Spectrometry (MS/MS)

The operation of two or more connected mass analyzers in sequence (or tandem) to perform one analysis is known as tandem mass spectrometry (MS/MS) [Busch et al., 1988; Hoffmann et al., 1996]. The basic tandem mass spectrometer is shown in Figure 2.3.

![Diagram of tandem mass spectrometry](image)

**Figure 2.3** A basic tandem mass spectrometer (MS/MS) system.

In MS/MS, a first mass analyzer (MS1) isolates or selects a precursor ion, $m_1^+$, which then undergoes fragmentations yielding product ions and neutral fragments:

$$m_1^+ \rightarrow m_2^+ + n$$

The parent ions with excess internal energy can undergo a specific or a number of competitive dissociations in an intermediate region or collision cell, yielding daughter ions $m_2^+$, $m_3^+$, $m_4^+$,.... , which are mass analyzed by the second mass analyzer (MS2). This is the most common daughter ion scan mode in tandem mass spectrometry, which is also used in the present study.
2.3.1 Metastable Ion Dissociation

Ions formed in an ion source of a mass spectrometer can be classified arbitrarily as unstable, stable and metastable ions depending on when or whether they fragment during their passage through the instrument. A hypothetical internal energy distribution for ions formed in an ion source of a magnetic sector type mass spectrometer is shown in Figure 2.4.

![Graph showing internal energy distribution](image)

**Figure 2.4** Hypothetical internal energy distribution, $P(E)$, versus $E$, the energy of precursor ions that lead to stable, metastable and unstable ions on the time scale of a typical magnetic sector type mass spectrometer.

Unstable ions are highly energetic ions that fragment within the ion source (time scale $10^{-6}$ s), and hence cannot be selected in the first mass analyzer (MS1) of tandem mass spectrometry. Stable ions are those that have lifetime longer than the time of passage through the flight path of the instrument to the electron multiplier detector (time scale $\sim 10^{-5}$ s). These ions can be mass selected by the MS1, but do
not have enough internal energy to undergo dissociation reactions. Metastable ions have intermediate internal energies, and dissociate outside of the ion source but before the detector. The metastable ions that are formed in the first field-free region (FFR-1, between the ionization source and the magnetic sector) of a magnetic (B) and electric (E) sector type mass spectrometer can be detected by \( B/E = \text{constant} \) linked-scans. The metastable ions, which are formed in the second field-free region (FFR-2, between the magnetic sector and the electrostatic sector) can be detected by ramping the DC voltage on the electrostatic analyzer (ESA). This is known as mass-analyzed ion kinetic energy (MIKE) scans. Only B/E linked-scans were used in the present study, which is described in more detail in Section 3.1.1.

Metastable ions are characterized by a relatively narrow range of internal energies, usually only slightly above the threshold of dissociation. The consequence is that most of the observable metastable dissociation reactions are with low energy barriers. The relative intensities of metastable ions are low, in the range of 0.01 – 1 % of the main beam intensity.

2.3.2 Collision-Induced Dissociation (CID)

Collisional activation is the most common excitation method used for activating the stable and metastable ions in tandem mass spectrometry. In the collision process, ions are accelerated to a desired kinetic energy and collide with the neutral / inert target gas (collision gas such as helium and argon). Part of translational energy is converted into internal energy of the ion. Once the internal energy \( (E) \) exceeds the critical energy \( (e_o) \) of a dissociation reaction, collision-induced dissociations (CID)
would occur. The amount of kinetic energy converted depends on a number of factors, such as the collision velocity and the nature of collision gas. In fact, the maximum amount of kinetic energy convertible to internal energy in a single collision is given by

\[ E_{(\text{con})} = E_{(\text{lab})} \frac{m_t}{m_p + m_t} \]  

[2.4]

where \( E_{(\text{con})} \) is the maximum energy convertible to internal energy, \( E_{(\text{lab})} \) is the kinetic energy of parent ion in laboratory reference frame, \( m_p \) is the mass of the parent ion, and \( m_t \) is the mass of the target gas. This expression is also the kinetic energy of the parent ion in center-of-mass reference frame often quoted in the literature. The numeric value of \( E_{(\text{con})} \) is much less than that of \( E_{(\text{lab})} \), since \( m_t \) is usually much smaller than \( m_p \).

**Low-Energy Collision-Induced Dissociation.** In low-energy CID, ions generally are accelerated to less than 100 eV kinetic energy (laboratory reference frame), as often is the case in the operation of the triple quadrupole tandem mass spectrometer. Hence, the interaction time between the target gas and the precursor ions is relative long, typically \( \sim 10^{-14} \) s, which is in the same order of magnitude of the period of vibration modes in molecules. Therefore, the ion excitation is mostly via vibrational excitation in low-energy CID processes. While the energy deposition in a single collision is small, the CID process is usually conducted under multiple collisions conditions. Thus, the average internal energy deposition after collisional activation is considerable, about 1 – 10 eV, depending on the instrumental and experimental conditions [Vekey, 1996]. However, it is not attainable to obtain very high internal energy, since the time frame of multiple collision activation is so long that the parent ion would have dissociated once it acquires moderately high internal energy. As a
result, the distribution of internal energy deposited in low-energy CID is much narrower than that of high-energy CID, and usually only fragmentation requiring lower critical energies are observed [Busch et al., 1988].

High-Energy Collision-Induced Dissociation. High-energy CID is usually conducted with magnetic sector instruments only. The precursor ions are accelerated to a kinetic energy of several keV, and allowed to collide with collision gas either in the first or second field-free region of the mass spectrometer. The activation mechanism in the high-energy CID is most likely direct electronic excitation through a curve-crossing mechanism [Levensen and Schwarz, 1983; Todd and McLafferty, 1983]. The interaction time between the target gas and the precursor is in the order of $\sim 10^{-16} - 10^{-15}$ s. This is too short to induce vibrational mode excitation (typically $\sim 10^{-14}$ s), but is about the characteristic time for excitation to higher electronic energy levels. Since the number of collisions in the high-energy CID is very small (predominately single collision at $> 70\%$ main beam transmission), the 'average' internal energy deposition is only about $1 - 3$ eV [Vekey, 1996]. However, in contrast to low-energy CID, there is a significant proportion of parent ions with very high internal energies ($> 10$ eV). Therefore, compared with low-energy CID spectra, there are relatively more undissociated parent ions, but fragmentation requiring high critical energy can also be observed [Busch et al., 1988].
2.4 The Mass Spectrometric Kinetic Method

The mass spectrometric kinetic method, developed by Cooks and co-workers [Cooks et al., 1981; Cooks and Wong, 1998], has been used to determine the thermochemical properties, such as proton affinities, electron affinities, organic cation affinities and metal cation affinities of a wide range of compounds. In this study, the kinetic method was used to determine the Ag (I) cation affinities of alkylbenzenes (RBz), substituted naphthalenes (X-Nap), substituted indoles (X-Indole), and substituted phenols (X-PhOH).

![Energy Diagram](image)

**Figure 2.5** Schematic energy diagram of the dissociation of a proton or metal cation bound heterodimer \([L_1 + M + L_2]^+\).
2.4.1 The Standard Kinetic Method

The kinetic method is based on two competitive dissociations of a proton or metal cation (or Ag⁺ in this study) bound heterodimer \([L_1 + M + L_2]^+\) of two ligands \(L_1\) and \(L_2\) to their respective monomer complexes \([M + L_1]^+\) and \([M + L_2]^+\) (Reaction 2.5a and 2.5b):

\[
\begin{align*}
\text{(2.5a)} & \quad k_1 \quad [L_1 + M + L_2]^+ \rightarrow [M + L_1]^+ + L_2 \\
\text{(2.5b)} & \quad k_2 \quad [L_1 + M + L_2]^+ \rightarrow [M + L_2]^+ + L_1
\end{align*}
\]

where \(k_1\) and \(k_2\) are the corresponding rate constants of these two dissociation pathways with corresponding critical energies \(\varepsilon_1\) and \(\varepsilon_2\), respectively. The schematic energy diagram for the dissociation pathways of a proton or metal cation bound heterodimer \([L_1 + M + L_2]^+\), as in Reaction 2.5a and 2.5b, is shown in Figure 2.5.

Applying the canonical transition state theory, the ratio of the corresponding rate constants is given by Equation 2.6.

\[
\ln\left(\frac{k_1}{k_2}\right) \approx \ln\left(\frac{[M + L_1]^+}{[M + L_2]^+}\right) = \ln\left(\frac{Q_1^*}{Q_2^*}\right) + \left(\varepsilon_2 - \varepsilon_1\right) / RT_{\text{eff}} \quad [2.6]
\]

\[
= \ln\left(\frac{Q_1^*}{Q_2^*}\right) + \Delta \varepsilon / RT_{\text{eff}} \quad [2.7]
\]

where \(Q_1^*\) and \(Q_2^*\) are the partition functions of the two transition states of the competing reactions, \(\Delta \varepsilon = \varepsilon_2 - \varepsilon_1\) is the difference in critical energies for the two competing reactions, \(R\) is the ideal gas constant and \(T_{\text{eff}}\) is the ‘effective temperature’ of the heterodimer undergoing dissociation. In a short time frame of the ion dissociations (~ \(10^{-6}\) s), the natural logarithm of the relative rates of dissociation, \(\ln(k_1 / k_2)\), can be approximated to the experimentally measurable natural logarithm
of the ion intensity ratio, \( \ln([M + L_1]^+ / [M + L_2]^+) \) (Equation 2.6). Also, it should be noted that since the \( M^+ \) bound heterodimer ions are isolated molecular systems undergoing dissociations in vacuum, \( T_{\text{eff}} \) is not a thermodynamic-based temperature parameter. It is, in fact, an empirical term non-linearly relating to an internal energy distribution of the dissociating \( M^+ \) bound heterodimer complex [Vekey and Drahos, 1999; Ervin 2000; 2002].

If the two dissociation pathways have no reverse activation barrier, the term \( \Delta \epsilon \) can be approximated by \( \Delta(\Delta H_{M^+}) \), i.e., \( \Delta H_{[M + L_1]^+} - \Delta H_{[M + L_2]^+} \) where \( \Delta H_{[M + L_1]^+} \) and \( \Delta H_{[M + L_2]^+} \) are enthalpy change of the Reaction 2.5a and 2.5b (i.e. the relative \( M^+ \) affinities), and the term \( \ln(Q_1^* / Q_2^*) \) can be approximated by the difference in entropy change between Reaction 2.5a and 2.5b, i.e., \( -\Delta(\Delta S_{M^+})^{\text{app}}/R = (\Delta S_1 - \Delta S_2)/R \). Then Equation 2.6 becomes Equations 2.8 and 2.9 by drawing analogous form of the thermodynamic formula \( G = H - T\Delta S \).

\[
\ln(k_1 / k_2) \approx \ln([M + L_1]^+ / [M + L_2]^+) = \Delta(\Delta H_{M^+})/RT_{\text{eff}} - \Delta(\Delta S_{M^+})^{\text{app}}/R \tag{2.8}
\]

\[
\approx \Delta(\Delta G_{M^+})^{\text{app}}/RT_{\text{eff}} \tag{2.9}
\]

When the two ligands in the heterodimer are structurally similar and forming the same type of bond(s) with \( M^+ \), the partition functions \( Q_1^* \approx Q_2^* \) and so \( \Delta(\Delta S_{M^+})^{\text{app}} \) can be approximated to be negligible. Thus Equation 2.8 can be further simplified to:

\[
\ln(k_1 / k_2) \approx \ln([M + L_1]^+ / [M + L_2]^+) \approx \Delta(\Delta H_{M^+})/RT_{\text{eff}} \tag{2.10}
\]

\[
= \frac{\Delta H_{[M + L_1]^+} - \Delta H_{[M + L_2]^+}}{RT_{\text{eff}}}
\]

The relative \( M^+ \) ion affinities of \( L_1 \) and \( L_2 \) would now be proportional to the natural logarithm of the ion intensity ratio \( \ln([M + L_1]^+ / [M + L_2]^+) \) at a given \( T_{\text{eff}} \). This
standard kinetic method is usually expressed in the form of a relative affinity $\Delta(\Delta H)$
ladder.

To convert the natural logarithm ratios to relative affinity, $\Delta(\Delta H)$, the factor $1/RT_{eff}$
(and hence $T_{eff}$) in Equation 2.10 must be determined. The best approach is to have
two or more ligands (points) of known $M^+$ affinities within the ladder, then a linear
regression plot of cumulative $ln([M + L_1]^+ / [M + L_2]^+)$ ratios versus the
corresponding known $M^+$ affinities ($\Delta H_{M^+}$) should yield a straight line with slope
equal to $1/RT_{eff}$ (Equation 2.10). The cumulative logarithm ratios of the stair-steps
relative to a particular reference (usually assigned to be the lowest binding affinity in
the ladder) are obtained by summation of the natural logarithm ratios of adjacent
stair-steps. The relative affinities with reference to the reference point, $\Delta(\Delta H)$, can
be obtained by multiplying the cumulative logarithm ratios with $1/RT_{eff}$. The
correlation ($R^2$) would also reveal the consistency of the ladder if more than two
reference ligands of known affinities were used. The linear regression plot also
correlates the cumulative logarithm ratios to the absolute affinities of the ‘unknown’
ligands within the affinity ladder. Therefore, by including two or more reference
ligands of known affinities in the affinity ladder, the absolute binding affinities of a
group of ligands can be determined efficiently by the standard kinetic method. This
is the approach used in the present study.
2.4.2 The Extended Kinetic Method

The extended mass spectrometric kinetic method is an ‘extended’ version of the standard kinetic method, developed by Fenselau and co-workers [Cheng et al., 1993], and Wesdemiotis and co-workers [Cerda and Wesdemiotis, 1996; Cerda et al., 1998], to allow the application of kinetic method between structurally dissimilar ligands.

Considering the competitive unimolecular dissociations of \( M^+ \) bound heterodimers of an ‘unknown’ ligand \( L_n \) and a series of reference ligands, \( L_i \), of known \( M^+ \) affinities (\( i = 1, 2, 3, 4, \) etc.). For the reference ligands \( L_i \) are structurally similar to each other but dissimilar to \( L_n \), the \( \Delta(\Delta S_{M^+})^{app} \) term is generally non-zero but assumed to be constant. According to Equation 2.8, the equation becomes Equation [2.12] with the analogy with the thermodynamic formula \( G = H - T\Delta S \).

\[
\ln\left(\frac{k_n}{k_i}\right) \approx \ln\left(\frac{[M + L_n]^+}{[M + L_i]^+}\right)
= -\Delta(\Delta S_{M^+})^{app}/RT + \frac{\Delta H_{[M + L_n]^+} - \Delta H_{[M + L_i]^+}}{RT_{\text{eff}}}
= (\Delta G_{M^+})^{app}/RT_{\text{eff}} - \Delta H_{[M + L_i]^+}/RT_{\text{eff}}
\]

[2.11]

where \( (\Delta G_{M^+})^{app} = \Delta H_{[M + L_n]^+} - T_{\text{eff}} \Delta(\Delta S_{M^+})^{app} \)

or

\[
(\Delta G_{M^+})^{app}/RT_{\text{eff}} = \Delta H_{[M + L_i]^+}/RT_{\text{eff}} - \Delta(\Delta S_{M^+})^{app}/R
\]

[2.12]

where \( (\Delta G_{M^+})^{app} \) and \( \Delta(\Delta S_{M^+})^{app} \) are the apparent free energy change and apparent entropy change for Reaction 2.5a and 2.5b.

According to Equation 2.11, a plot of \( \ln\left(\frac{[M + L_n]^+}{[M + L_i]^+}\right) \) versus \( \Delta H_{[M + L_i]^+}/RT_{\text{eff}} \) would yield a straight line with slope and \( y \)-intercept corresponding to \(-1/RT_{\text{eff}}\) and \((\Delta G_{M^+})^{app}/RT_{\text{eff}}\), respectively. Furthermore, by determining several \( T_{\text{eff}}(s) \) (or
internal excitation energies), obtained by varying the collision energies or the collision gas pressure, a second plot of \((\Delta G_{M^+}^{app}/RT_{eff})\) versus \(1/RT_{eff}\) gives a straight line with slope and y-intercept equal to \(\Delta H_{[M^+ + L_n]^+} + (\Delta S_{M^+}^{app}/R)\), respectively (Equation [2.12]) and therefore, the \(M^+\) affinity of the 'unknown' ligand \(L_n\), \(\Delta H_{[M^+ + L_n]^+}\) can be obtained.

However, the \((\Delta G_{M^+}^{app})\) and the \(1/RT_{eff}\) terms are, in fact, covariant so that the highly linear \((\Delta G_{M^+}^{app}/RT_{eff})\) versus \(1/RT_{eff}\) in the second plot in the original extended kinetic method could be an artifact in data analysis [Armentrout, 2000]. To remove this possible artifact, Armentrout suggested a statistical data treatment approach: by introducing the \(\Delta H_{Avg}\) term, where \(\Delta H_{Avg}\) is the average of the \(\Delta H_{[M^+ + L_i]^+}\) values of the reference ligands (\(L_i\)). The first plot now becomes a plot of \(ln(\left [ [M + L_n]^{+} / [M + L_i]^{+} \right])\) against \(\Delta H_{[M^+ + L_n]^+} + \Delta H_{Avg}\), with the slope \(m = -1/RT_{eff}\) and the intercept \(y_0 = ((\Delta G_{M^+}^{app} - \Delta H_{Avg})/RT_{eff})\). This is followed by a second plot of \(y_0\) versus \(-m\) which yields a slope of \(\Delta H_{[M^+ + L_n]^+} + \Delta H_{Avg}\), and an y-intercept of \(\Delta(\Delta S_{M^+}^{app}/R)\), from which the unknown \(\Delta H_{[M^+ + L_n]^+}\) and \(\Delta(\Delta S_{M^+}^{app})\) terms can be obtained. This statistical approach as proposed by Armentrout was adopted in the data treatment process in the present study.

In deriving the equations used in the protocols of the mass spectrometric kinetic method, one of the most fundamental assumption is the application of transition state theory, which assumes that the systems are in thermodynamic equilibrium (Boltzmann distribution). A system of isolated ions in mass spectrometer is not describable by a thermodynamic temperature. However, alternative derivations which do not invoke the transition state theory and thermodynamic equilibrium have
lead to the same linear free-energy type of relationships used in the protocols of the kinetic method [Bojesen and Breindahl, 1994; Grutzmacher and Caltapanides, 1994; Craig et al., 1997]. These studies provide the physico-chemical basis of the kinetic method. It should be pointed out that the $T_{\text{eff}}$, $(\Delta G_{\text{M}^+})^{\text{app}}$ and $(\Delta S_{\text{M}^+})^{\text{app}}$ terms are related to the properties of the dissociating heterodimer, and are dependent on the instrumental conditions under which the dissociation reactions are monitored. The theoretical basis, assumptions and the methodology of the kinetic method have been critically reviewed and discussed recently [Cooks and Wong, 1998; Cooks et al., 1999; Armentrout, 1999; Drahos and Vekey, 1999; Ervin, 2000; 2002]. The major advantage of the kinetic method for determining thermochemical properties is its simplicity, and the measurements can be carried out in commonly available tandem mass spectrometers.
Chapter 3  Instrumentation and Experimental Methodology

3.1  Mass Spectrometers

All experimental studies were carried out by a Finnigan-MAT 95S B-E tandem mass spectrometer (Bremen, Germany), a Micromass VG Quattro II triple quadrupole tandem mass spectrometer (Manchester, UK) and a Micromass Quattro Ultima triple quadrupole tandem mass spectrometer (Manchester, UK). All the mass spectrometers were installed at the Department of Applied Biology and Chemical Technology, The Hong Kong Polytechnic University, except the Micromass VG Quattro II triple quadrupole tandem mass spectrometer, which is by the courtesy of the Department of Chemical Pathology, The Chinese University of Hong Kong.

3.1.1  Finnigan-MAT 95S B-E Tandem Mass Spectrometer System

*General Configuration.* The Finnigan-MAT 95S tandem mass spectrometer is B-E sector double-focusing tandem mass spectrometer (MS1 in Figure 3.1). It includes a 65°, 35 cm radius magnetic sector, B, followed by an electric sector, E. The full accelerating voltage is 5 kV (but is usually set at 4.7 kV) and the maximum value of magnetic field strength is 1.7 Tesla. At full accelerating voltage, the mass spectrometer has a mass range of 3,500 Da and a maximum resolution of 60,000 (10% valley) in the double focusing mode.
Figure 3.1  Schematic diagram of the Finnigan-MAT 95ST B-E (MS1) ion trap (MS2) tandem mass spectrometer system.

The region located between the ionization source and the first analyzer (the magnetic sector B with radius $r_B$) is called the first field-free region (FFR-1), while the second field-free region (FFR-2) is located between the first magnetic sector and the second electric sector. A collision cell is mounted within the first field free region (FFR-1) between the ionization source and the magnetic sector where collision gas (e.g. helium, argon or nitrogen) is introduced. The vacuum at the ion source ($2.5 \times 10^{-8}$ mbar) and analyzer ($2 \times 6 \times 10^{-8}$ mbar) regions as monitored by ionization gauges, and are maintained by two separate turbo-molecular pumps, and backed by two rotary pumps.

The thermally assisted electrospray ionization (ESI) interface is an add-on module (Figure 3.2) [Bruins, 1997]. An extra pumping system (since the ESI source is opened to atmosphere) is provided by an additional rotary pump and a turbo-
molecular pump to maintain the vacuum at $3 - 6 \times 10^{-5}$ mbar at the ESI source head. The ion beam is detected by an off-axis convection dynode ($\pm 20$ kV) - Cu/Be electron multiplier detector system (maximum up to $\pm 2$ kV). Data acquisition and instrumental control were carried out with the aid of the Finnigan-MAT ICIS system software and a DEC Alpha Workstation.

![Diagram](image)

**Figure 3.2** Schematic diagram of the thermally assisted electrospray source interface of the Finnigan MAT 95ST mass spectrometer system.

**Magnetic Sector (B) Mass Analyzer.** A magnetic sector acts a momentum-focusing element for charged ions. When an ion with mass $m$ and charge $ze$ generated from the ionization source enter to a magnetic sector, the ion will follow in a circular trajectory in a plane normal to the direction of the magnetic field. The motion of the ion in the magnetic field is basically operated by Equation 3.1.

$$\frac{m}{z} = \frac{B^2 r^2 e}{2V} \quad [3.1]$$

where $m = \text{mass of the ion (in kilograms)},$

$z = \text{charge on the ion (in Coulombs)},$

$B = \text{magnetic field strength (in Tesla)},$
\( r_B \) = radius of the magnetic sector (in meters),

\( e \) = charge of an electron \((1.6 \times 10^{-19} \text{ Coulombs})\),

\( V \) = accelerating voltage (in volts).

At a given magnetic field strength with constant accelerating voltage, ions of only one mass-to-charge (m/z) ratio will follow the trajectory along the central radius of the magnetic sector. The m/z ratio of the ion focused along this trajectory is thus a function of magnetic field strength and accelerating voltage, either of this can be varied to change the mass of ion that follows this trajectory. Ions of other mass-to-charge ratios will be lost in collision with the wall of the magnetic sector from different radii of trajectories. In the Finnigan-MAT 95S magnetic sector mass spectrometer, a normal spectrum is obtained by varying the magnetic field strength B.

**Electric Sector (E) Mass Analyzer.** The electrostatic analyzer selects ions according to their kinetic energy \((mv^2/2)\) but not mass. For an electrostatic analyzer (ESA) with a radical electric field, \( E \), and radius of curvature, \( r_E \), the travelling ions experience an electrostatic force governed by Equation 3.2.

\[
E ze = \frac{mv^2}{r_E} \quad \text{or} \quad mv^2 = E z e r_E \quad [3.2]
\]

Combining Equation 3.1 and 3.2 leads to 3.3. At a fixed and constant \( E \), only ions (after passing through B sector) with certain kinetic energy can pass through the electrostatic sector and are detected.

\[
E = \frac{mv^2}{z e r_E} = \frac{2V}{r_E} \quad [3.3]
\]

where \( v \) = velocity of the ion (in meters/second),
E = electric field strength (in volts/meter),
\(r_E\) = radius of the deflection in the electric sector (in meters).

**B/E linked Scans.** Fragment ions arisen from unimolecular dissociation(s) occurring in the FFR-1 of a double-focusing B-E sector type mass spectrometer can be observed by a B/E linked scan, which consists of a simultaneous scan of the B and E sectors according to a mathematical relationship dependent on the system B-E geometry. For a fragmentation reaction occurring in the first field free region (FFR-1) between the ionization source and the magnetic sector,

\[ m_1^+ \rightarrow m_2^+ + n \]

the parent ion, \(m_1^+\), enters the FFR-1 with the mass, \(m_1\), and velocity, \(v_1\), possesses the kinetic energy given by Equation 3.4.

\[ \text{K.E.} = \frac{1}{2} m_1 v_1^2 = \text{eV} \]  \[3.4\]

By the law of conservation of momentum, \(m_1 v_1 = (m_2 + n)v_1\), the velocity of the daughter ion, \(m_2^+\), will be the same as that of parent ion, \(m_1^+\). Thus, the magnetic field (B) and electric field (E) required to transmit the \(m_1^+\) and \(m_2^+\) are:

\[ B_1 = \frac{m_1 v_1}{r_B e} \quad \text{and} \quad B_2 = \frac{m_2 v_1}{r_B e} \]  \[3.5\]

and

\[ E_1 = \frac{m_1 v_1^2}{r_E e} \quad \text{and} \quad E_2 = \frac{m_2 v_1^2}{r_E e} \]  \[3.6\]

where \(r_B\) and \(r_E\) are the radii of the ion path in the magnetic sector and electric sector, respectively. By combining Equation 3.5 and 3.6, the equation can be simplified as Equation 3.7.
\[
\frac{B_1}{E_1} = \frac{B_2}{E_2} = \frac{r_E}{r_B} \nu_1 = \text{constant} \tag{3.7}
\]

Therefore, if \(B\) and \(E\) are set to transmit the parent ion, \(m_1^+\), and the values of \(B\) and \(E\) are simultaneously varied while keeping \(B/E\) constant, all its daughter ions, \(m_2^+\) will be transmitted subsequently. Thus, a \(B/E\) linked-scan will yield a daughter ion spectrum of \(m_1^+\) decomposing in the FFR-1 of a double-focusing B-E mass spectrometer. \(B/E\) linked scans were used extensively in our present study.

Normally, a metastable ion spectrum resulting from unimolecular dissociation reactions in the FFR-1 is obtained with \(B/E\) linked scan in the absence of collision gas. With introduction of a collision gas (usually helium used in this study), the precursor ion beam, \(m_1^+\), (accelerated at a source potential of 4.7 kV) will undergo high-energy CID. However, in the absence of collision gas in ESI operation, a high-energy CID spectrum, resulted from the collision with the residual solvent (methanol) vapor in the collision cell adjacent to the ESI source, can also be obtained.
3.1.2 Micromass VG Quattro II Triple Quadrupole Tandem Mass Spectrometer System

**General Configuration.** The Micromass Quattro II triple quadrupole tandem mass spectrometer (Figure 3.3), consists of two quadrupole mass analyzers (MS1 and MS2), and a hexapole collision cell situated in-between. The first quadrupole (MS1) provides either mass scanning (as single quadrupole mass analyzer) or mass filtered the selected precursor ions for CID analysis. The hexapole collision cell is a radio-frequency (RF) only device. Collision gas (argon) is introduced in the collision cell where ions passing through MS1 are then collisionally activated and undergo fragmentations. The RF field within the hexapole collision cell efficiently refocuses ions scattered by the CID process. Ions emerge from the collision cell are transmitted to the second quadrupole (MS2) for mass analysis and subsequently detected by a Dynolite detector system (phosphor and photomultiplier). The vacuum at source region is maintained at about $1.5 \times 10^{-2}$ mbar under ESI mode and the analyzer region is kept at about $2.5 \times 10^{-5}$ mbar (as being monitored by active inverted magnetron gauges and Pirani gauges). Data acquisition and instrumental control were carried by the MassLynx NT software system.

![Figure 3.3 Schematic diagram of the Micromass VG Quattro II triple quadrupole mass spectrometer system.](image)

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The Micromass Quattro II triple quadrupole tandem mass spectrometer is equipped with a pneumatically assisted electrospray interface (Figure 3.4) where a stream of nebulizing gas (N₂) and drying gas (N₂) at slightly evaluated ESI chamber (typically set about 60 – 80°C) facilitate charged droplets formation, solvent evaporation, and de-clustering of solvent cluster ions.

Figure 3.4 Schematic diagram of the pneumatically assisted electrospray source interface of the Micromass VG Quattro II mass spectrometer system.
3.1.3 Micromass Quattro Ultima Triple Quadrupole Tandem Mass Spectrometer System

General Configuration. The Micromass Quattro Ultima triple quadrupole tandem mass spectrometer (Figure 3.5) is a modified version of the Micromass VG Quattro II triple quadrupole mass spectrometer aforementioned, but with a Z-Spray ESI source (Figure 3.6). The configuration of the mass spectrometer, instrumental control and operation are referred to those of VG Quattro II instrument.

![Schematic diagram of the Micromass Quattro Ultima mass spectrometer system.](image)

A standard Micromass Z-spray ESI source is a modified pneumatically assisted electrospray ionization interface, which is shown in Figure 3.6. Its design feature is
to facilitate the transport of ionic species to the mass analyzer orthogonally and at
the same time greatly reducing the buildup of deposits and blockages of the skimmer
orifice. This allows a higher sensitivity and stable performance of the ESI process
over a long period of operating time.

Figure 3.6 Schematic diagram of the pneumatically assisted Z-spray
electrospray source interface of the Micromass Quattro Ultima
mass spectrometer system.
### 3.2 Chemicals and Sample Preparation

#### 3.2.1 Chemicals

Analytical grade chemicals were purchased from Acros, Aldrich, Sigma, Chem Service and Lancaster chemicals companies, and they were used as received. The chemicals and reagents used in the present study are listed below.

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<th>Abbreviation</th>
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### Substituted Naphthalenes

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<td>4-Methylphenol</td>
<td>Me-PhOH</td>
<td>108</td>
</tr>
<tr>
<td>4-Ethylphenol</td>
<td>Et-PhOH</td>
<td>122</td>
</tr>
<tr>
<td>4-i-Propylphenol</td>
<td>i-Pr-PhOH</td>
<td>136</td>
</tr>
<tr>
<td>4-t-Butylphenol</td>
<td>t-Bu-PhOH</td>
<td>150</td>
</tr>
<tr>
<td>4-Methoxyphenol</td>
<td>MeO-PhOH</td>
<td>124</td>
</tr>
<tr>
<td>4-Ethoxyphenol</td>
<td>EtO-PhOH</td>
<td>138</td>
</tr>
<tr>
<td>4-Nitrophenol</td>
<td>NO2-PhOH</td>
<td>139</td>
</tr>
</tbody>
</table>
3.2.2 Sample Preparation

Stock solutions of $2 \times 10^{-3}$ M AgNO$_3$, 1.0 M alkylbenzenes, 0.1 M substituted naphthalenes, 0.1 M substituted indoles, 1.0 M substituted phenols were prepared in methanol individually. $1 \times 10^{-2}$ M and $1 \times 10^{-3}$ M substituted naphthalenes and substituted indoles were also prepared by serial dilution of the stock solutions. The Ag$^+$ bound heterodimer complexes were generated by mixing two sample solutions of appropriate concentrations and 10 – 100 uL AgNO$_3$ solution in 1mL methanol. The concentrations of ligands in the mixture solution are not critical, but were optimized to get a maximum ion intensity signal of the Ag$^+$ bound heterodimer. The solutions were introduced into the ESI interface by a syringe pump (Harvard Apparatus, model 22) at a flow rate of 5 – 10 μL/min.
3.3 Instrumentation Conditions

3.3.1 High-Energy CID Studies in Finnigan-MAT 95S B-E Tandem Mass Spectrometer

Under He-CID conditions, helium gas was introduced to the collision cell at FRR-1 to yield 30 – 40 % attenuation of the parent ion beam. The key instrumental parameters are listed below.

**Ion Generation – ESI Tuning Parameters:**

- Ionization Mode : ESI positive
- Electrospray Capillary Voltage : 2.2 – 2.4 kV
- Heated Capillary Temperature : 190 – 220 °C
- Heated Capillary Voltage : + 4 – 35 V
- Tube lens Voltage : + 20 – 60 V
- Skimmer Lens Voltage : + 2 – 6 V
- Octapole Voltage : + 0.3 – 0.9 V
- Nebulizing Gas : Nil

**Mass Analyzer Operating Parameters:**

- Scan Type : Constant B/E linked scan
- Accelerating Voltage : 4.7 kV
- Multiplier : 1.8 kV
- Conversion Dynode : 15 kV
- Scan Rate : 10 – 20 s/decade
- Interscan Time : 0.4 s
Pressure Conditions Recorded from the Ion Gauge Reading:

Ion Source Vacuum : \( 6 - 7 \times 10^{-5} \) mbar
Electrostatic Analyzer Vacuum : \( 4 - 6 \times 10^{-8} \) mbar

Low resolution profile scans were acquired. Typically 60 – 100 scans were accumulated to yield a mass spectrum. The ion intensity (peak height) ratios, \([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+\), taken as an average of 3 – 5 runs, are reproducible to within \( \pm 7 \% \) (coefficient of variation, \( n = 3 – 5 \)).

3.3.2 Low-Energy CID Studies in Micromass VG Quattro II Triple Quadrupole Tandem Mass Spectrometer

The validation of experimental absolute \( \text{Ag}^+ \) affinities of alkylbenzenes by extended kinetic method were performed in the Micromass VG Quattro II Triple Quadrupole Tandem Mass Spectrometer. Typical ion generation and mass spectrum acquisition conditions are shown as below.

Ion Generation – ESI Tuning Parameters:

Ionization Mode : ESI positive
Electrospray Capillary : 2.8 kV
HV Lens : 0.05 kV
Cone Voltage : 10 V
Skimmer Lens Offset : 5 V
Source Temperature : 60 °C
Flow Rate of Nebulizing Gas (\( \text{N}_2 \)) : 10 L/hr
Flow Rate of Drying Gas (\( \text{N}_2 \)) : 250 L/hr
**Mass Analyzer (MS1) Operating Parameters:**

- Low Mass (LM) Resolution Setting : 15.0
- High Mass (HM) Resolution Setting : 15.0
- Ion Energy Ramp : 2.0 V
- Lens 6 : 2 V

**Mass Analyzer (MS2) Operating Parameters:**

- Scan Mode : Daughter scan (MS/MS)
- Low Mass (LM) Resolution Setting : 15.0
- High Mass (HM) Resolution Setting : 15.0
- Ion Energy : 1.0 V
- Ion Energy Ramp : 0.0 V
- Collision : 10 – 50 eV
- Lens 8 : 40 V
- Lens 9 : 0 V
- Multiplier : 650 V
- Scan Rate : 300 – 400 Th/s
- Interscan Time : 0.2 s
- Collision Gas (Collision Cell) : Argon

**Pressure Conditions Recorded from the Magnetron (Penning) Gauge Reading:**

- Inlet Vacuum : 1.2 – 1.4 x 10^{-2} mbar
- Analyzer Vacuum : 1.2 – 2.7 x 10^{-5} mbar
- Collision Cell : 2.1 – 5.8 x 10^{-4} mbar

Typically 30 – 40 scans were accumulated to yield a mass spectrum. The ion intensity (peak height) ratios, [Ag + L1]^+ / [Ag + L2]^+, taken as an average of 3 – 4 runs, are reproducible to within ± 5% (coefficient of variation, n = 3 – 4).
3.3.3 Low-Energy CID Studies in Micromass Quattro Ultima Triple Quadrupole Tandem Mass Spectrometer

Most of the low-energy CID measurements were conducted in the Micromass Quattro Ultima Triple Quadrupole Tandem Mass Spectrometer. The collision gas was introduced to yield 20 – 30 % attenuation of the parent ion beam. The typical operating key parameters are shown as below.

*Ion Generation – ESI Tuning Parameters:*

- **Ionization Mode:** ESI positive
- **Electrospray Capillary:** 3.0 – 3.8 kV
- **Cone Voltage:** 10 – 25 V
- **Hexapole 1:** 0.0 – 0.2 V
- **Aperture:** 0.0 – 0.2 V
- **Hexapole 2:** 0.0 – 0.2 V
- **Source Temperature:** 70 – 80 °C
- **Desolvation Temperature:** 120 – 150 °C
- **Nebulizing Gas (N₂):** Fully closed
- **Flow Rate of Desolvation Gas (N₂):** 400 – 600 L/hr
- **Flow Rate of Cone Gas (N₂):** 40 – 70 L/hr

*Mass Analyzer (MSI) Operating Parameters:*

- **Low Mass (LM) Resolution Setting:** 15.0
- **High Mass (HM) Resolution Setting:** 15.0
- **Ion Energy:** 0.1 – 0.3 V
**Mass Analyzer (MS2) Operating Parameters:**

- **Scan Mode**: Daughter scan (MS/MS)
- **Low Mass (LM) Resolution Setting**: 15.0
- **High Mass (HM) Resolution Setting**: 15.0
- **Ion Energy**: 0.1 – 0.3 V
- **Ion Energy Ramp**: 0.0 V
- **Entrance**: 1 V
- **Exit**: 1 V
- **Collision**: 10 – 50 eV
- **Multiplier**: 650 V
- **Scan Rate**: 350 – 450 Th/s
- **Interscan Time**: 0.1 s
- **Collision Gas (Collision Cell)**: Argon

**Pressure Conditions Recorded from the Ion Gauge Reading:**

- **Analyzer (Penning) Vacuum**: $1.70 - 2.15 \times 10^{-5}$ mbar
- **Collision Cell (Pirani)**: $< 1 \times 10^{-4}$ mbar

Typically 60 – 80 scans were accumulated to yield a mass spectrum. The ion intensity (peak height) ratios, $[\text{Ag} + \text{L}_1]^+ / [\text{Ag} + \text{L}_2]^+$, taken as an average of 4 – 5 runs, are reproducible to within ± 7 – 8 % (coefficient of variation, n = 4 – 5).
Chapter 4 Ag (I) Cation Affinities of Alkylbenzenes

4.1 Background

Benzene is the simplest aromatic-π model system that could mimic the binding properties of cation-π synthetic 'host' ligands such as calixarenes and cyclophanes [Ma and Dougherty, 1997]. A fundamental understanding of the origin of cation-π interactions in benzene, therefore, is useful in understanding the cation-π interactions in more complicated aromatic systems.

Earlier gas-phase experimental studies showed that the binding of Li⁺ [Woodin and Beauchamp, 1978], Na⁺ [Guo et al., 1990; Armentrout and Rodgers, 2000] and K⁺ [Sunner et al., 1981] to benzene is indeed strong and comparable to the bonding strengths of more typical cation coordinating functional groups, such as water, alcohols, and ethers. Recently, the alkali metal cation affinities (Li⁺, Na⁺, K⁺, Rb⁺ and Cs⁺) of benzene [Amicangelo and Armentrout, 2000], toluene [Ammunugama and Rodgers, 2002a], fluorobenzenes [Ammunugama and Rodgers, 2002b], phenol [Ammunugama and Rodgers, 2002c], methoxybenzene [Ammunugama and Rodgers, 2003c] and aniline [Ammunugama and Rodgers, 2003b] have been reported. However, experimental studies investigating the cationic characteristics of transition metals, such as Ag⁺ in cation-π interaction, are very limited.

This chapter describes a systematic study on the intrinsic binding affinity of Ag⁺ to alkyl and multiple-methyl substituted benzenes. For simplicity of description and unless otherwise explicitly stated, 'alkylbenzenes (RBz)' is used here in the broad sense to denote both alkyl and multiple-methyl substituted benzenes. Using the
theoretical Ag\(^+\) affinities of toluene, m-xylene, 1,3,5-trimethylbenzene and pentamethylbenzene as reference values, the relative and absolute Ag\(^+\) affinities of alkylbenzenes were determined by the standard kinetic method. The affinity values so obtained were then used to determine the Ag\(^+\) affinities of acetone and acetonitrile, which in turn were compared with literature values. We will show later that our kinetic method Ag\(^+\) affinity values for acetone and acetonitrile are in very good agreement (within ± 8 kJ mol\(^{-1}\)) with experimental values found in literature, thus validating the experimental Ag\(^+\) affinities of alkylbenzenes obtained in this study.

4.2 Results and Discussion

4.2.1 Determination of Ag (I) Cation Affinities of Alkylbenzenes

Relative Ag (I) Cation Affinities of Alkylbenzenes. We first attempted to establish the order of Ag\(^+\) affinities among alkylbenzenes by the standard kinetic method. Silver has two isotopes (\(^{107}\text{Ag}\) and \(^{109}\text{Ag}\)) of near equal abundance; the mass spectra of Ag\(^+\)-ligand complexes usually display isotopic Ag\(^+\) adduct ions of almost equal intensities / heights. The \(^{107}\text{Ag}\) adduct ions were usually chosen for CID studies in the present work. For simplicity of description, \(^{107}\text{Ag}\) is simply to Ag hereafter. The ESI generated Ag\(^+\) bound heterodimers of alkylbenzenes, [RBz\(_1\) + Ag + RBz\(_2\)]\(^+\) (RBz\(_1\) and RBz\(_2\) are two different alkylbenzenes), were selected as the precursor ions and were allowed to dissociate to its [Ag + RBz\(_1\)]\(^+\) and [Ag + RBz\(_2\)]\(^+\) monomer complexes in the collision cell of a Micromass Quattro Ultima triple quadrupole tandem mass spectrometer. A typical low-energy (25 eV, laboratory frame) Ar-CID mass spectrum of Ag\(^+\) bound heterodimer between m-xylene (m-Xy) and 1,3,5-
trimethylbenzene (1,3,5-Me₂Bz), \([(m-Xy) + Ag + (1,3,5-Me₂Bz)]^+\), is shown in Figure 4.1. The heterodimer dissociates by the competitive loss of \(m\)-Xy and 1,3,5-Me₂Bz to yield fragment ions at \(m/z\) 227 and \(m/z\) 213, respectively. The only two predominant decomposition channels observed in the CID mass spectrum suggest that the two alkylbenzenes in the heterodimer complex are loosely bound to Ag⁺.

![Graph](image)

**Figure 4.1** Low-energy (25eV, laboratory frame) Ar-CID mass spectrum of Ag (I) cation bound heterodimer, \([(m-Xy) + Ag + (1,3,5-Me₂Bz)]^+\), \((m-Xy = m\)-xylene and 1,3,5-Me₂Bz = 1,3,5-trimethylbenzene).

As shown in Figure 4.1, the ion intensity of \([Ag + (1,3,5-Me₂Bz)]^+\) is higher than that of \([Ag + (m-Xy)]^+\), suggesting that 1,3,5-Me₂Bz has a higher Ag⁺ binding affinity than \(m\)-Xy. The ion intensity ratio, \([Ag + (1,3,5-Me₂Bz)]^+ / [Ag + (m-Xy)]^+\), is 2.74 \((ln([Ag + (1,3,5-Me₂Bz)]^+ / [Ag + (m-Xy)]^+ = 1.01)\). According to Equation 2.10, the experimental determined \(ln([Ag + (1,3,5-Me₂Bz)]^+ / [Ag + (m-Xy)]^+\) value is directly proportional to the relative Ag⁺ affinity, i.e. \(\Delta H_{[Ag + (1,3,5-Me₂Bz)]^+ - [Ag + (m-Xy)]^+}\).
<table>
<thead>
<tr>
<th>Alkylbenzene, RBz</th>
<th>( \ln([\text{Ag} + \text{RBz}_1]^+ /[\text{Ag} + \text{RBz}_2]^+) )</th>
<th>( \ln([\text{Ag} + \text{Bz}]^+ /[\text{Ag} + \text{Bz}]^+) )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bz</td>
<td></td>
<td>0.00 ± 0.00</td>
</tr>
<tr>
<td>Tol</td>
<td>1.63</td>
<td>1.13 ± 0.03</td>
</tr>
<tr>
<td>EtBz</td>
<td>2.25</td>
<td>1.74 ± 0.06</td>
</tr>
<tr>
<td>m-Xy</td>
<td>1.22</td>
<td>1.74 ± 0.06</td>
</tr>
<tr>
<td>n-PrBz</td>
<td>0.90</td>
<td>2.26 ± 0.07</td>
</tr>
<tr>
<td>i-PrBz</td>
<td>0.41</td>
<td>2.32 ± 0.07</td>
</tr>
<tr>
<td>n-BuBz</td>
<td>1.01</td>
<td>2.36 ± 0.07</td>
</tr>
<tr>
<td>i-BuBz</td>
<td>0.65</td>
<td>2.73 ± 0.08</td>
</tr>
<tr>
<td>s-BuBz</td>
<td>0.59</td>
<td>2.78 ± 0.08</td>
</tr>
<tr>
<td>t-BuBz</td>
<td>0.49</td>
<td>2.87 ± 0.09</td>
</tr>
<tr>
<td>1,3,5-Me₃Bz</td>
<td>0.56</td>
<td>3.01 ± 0.10</td>
</tr>
<tr>
<td>n-PnBz</td>
<td>0.88</td>
<td>3.33 ± 0.10</td>
</tr>
<tr>
<td>1,2,4,5-Me₄Bz</td>
<td>2.53</td>
<td>3.44 ± 0.11</td>
</tr>
<tr>
<td>Me₅Bz</td>
<td>2.31</td>
<td>4.12 ± 0.11</td>
</tr>
<tr>
<td>Me₆Bz</td>
<td></td>
<td>5.10 ± 0.11</td>
</tr>
</tbody>
</table>

**Figure 4.2** Experimentally measured \( \ln([\text{Ag} + \text{RBz}_1]^+ /[\text{Ag} + \text{RBz}_2]^+) \) values for the low-energy (25 eV, laboratory frame) Ar-CID of \( \text{Ag}^+ \) bound heterodimers of alkylbenzenes (\( T_{\text{eff}} = 1,241 \) K). The \( \ln([\text{Ag} + \text{RBz}_1]^+ /[\text{Ag} + \text{RBz}_2]^+) \) values are the logarithm of ion intensity ratios. The data presented under the heading \( \ln([\text{Ag} + \text{RBz}]^+ /[\text{Ag} + \text{Bz}]^+) \) are average ± standard deviation of cumulative values expressed relative to benzene (Bz). The values in ( ) are the arithmetic difference of measured values having a common reference alkylbenzene.

Altogether, the decomposition of thirty \( \text{Ag}^+ \) bound heterodimers of alkylbenzenes were determined under low-energy Ar-CID conditions. The \( \ln([\text{Ag} + \text{RBz}_1]^+ /[\text{Ag} \text{Bz}]^+) \) values...
values are presented in the relative Ag⁺ affinity ladder (at T_{eff} = 1,241 K) shown in Figure 4.2. The values under the heading ln([Ag + RBz]⁺ / [Ag + Bz]⁺) are the cumulative logarithm ratios relative to benzene (Bz). Alkylbenzenes with higher Ag⁺ affinities are placed at the bottom of the ladder and those with lower Ag⁺ affinities are placed at the top of the ladder. The ln([Ag + RBz₁]⁺ / [Ag + RBz₂]⁺) values of isobaric alkylbenzenes (same molecular mass) such as EtBz and m-Xy, cannot be measured directly, but can be determined from their respective heterodimers with other alkylbenzenes. For example, the ln([Ag + (m-Xy)]⁺ / [Ag + Tol]⁺) and ln([Ag + EtBz]⁺ / [Ag + Tol]⁺) values are 1.16 and 0.61, respectively; their arithmetic difference, corresponding to ln([Ag + (m-Xy)]⁺ / [Ag + EtBz]⁺), is 0.55. Similar treatment applies to the ln([Ag + (n-PrBz)]⁺ / [Ag + EtBz]⁺) (0.55) and ln([Ag + (n-PrBz)]⁺ / [Ag + (m-Xy)]⁺) (0.06), and the difference yields a ln([Ag + (m-Xy)]⁺ / [Ag + EtBz]⁺) value of 0.49. The average of 0.55 and 0.49 is 0.52, and is taken to be the ln([Ag + (m-Xy)]⁺ / [Ag + EtBz]⁺) value.

As shown in Figure 4.2, the 'internal consistency' among alkylbenzenes is observed in the 'additivity' comparison of the stair-step ln([Ag + RBz₁]⁺ / [Ag + RBz₂]⁺) values. For example, the ln([Ag + Me₅Bz]⁺ / [Ag + (1,3,5-Me₃Bz)]⁺) value for the [Me₅Bz + Ag + (1,3,5-Me₃Bz)]⁺ heterodimer is 2.53 from direct measurement, which is close to the value of 2.44 (i.e. 0.11 + 0.68 + 0.98 + 0.67) obtained from the summation of five successive stair-step ln([Ag + RBz₁]⁺ / [Ag + RBz₂]⁺) values from 1,3,5-Me₃Bz to Me₅Bz (Figure 4.2). The difference is 0.09, which is comparable to the experimental uncertainty (largest SD is ± 0.05 in direct intensity ratio). The results suggest that entropic effects in the dissociation of the Ag⁺ bound heterodimers of alkylbenzenes are indeed very small or negligible, and Equation 2.10 is a valid approximation under the low-energy CID conditions. Therefore, the
\ln([\text{Ag} + \text{RBz}_1]^+ / [\text{Ag} + \text{RBz}_2]^+) \) ratios can be approximated to the \( \Delta(\Delta H_{\text{Ag}}^+) / RT_{\text{eff}} \) term, where \( \Delta(\Delta H_{\text{Ag}}^+) \) is the relative Ag\(^+\) affinity between the alkylbenzenes RBz\(_1\) and RBz\(_2\). Thus, the relative order of Ag\(^+\) affinities of fifteen alkylbenzenes, expressed in terms of the cumulative \( \ln([\text{Ag} + \text{RBz}]^+ / [\text{Ag} + \text{Bz}]^+) \) values, were found in the order:

- benzene (Bz) < toluene (Tol) < ethylbenzene (EtBz) < \text{-}xylene (\text{-}Xy) < n-propylbenzene (n-PrBz) < i-propylbenzene (i-PrBz) < n-butylbenzene (n-BuBz) < i-butylbenzene (i-BuBz) < s-butylbenzene (s-BuBz) < t-butylbenzene (t-BuBz) < 1,3,5-trimethylbenzene (1,3,5-Me\(_3\)Bz) < n-pentylbenzene (n-PnBz) < 1,2,4,5-tetramethylbenzene (1,2,4,5-Me\(_4\)Bz) < pentamethylbenzene (Me\(_5\)Bz) < hexamethylbenzene (Me\(_6\)Bz)

To confirm the order of the Ag\(^+\) affinities obtained under low-energy (eV, laboratory frame / scale) Ar-CID conditions, independent kinetic method measurements at higher collisional excitation energy (by high-energy He-CID) were conducted. Under higher ‘effective temperature, \( T_{\text{eff}} \)’ CID conditions, any entropic effects in the dissociation of the Ag\(^+\) bound heterodimers (i.e. the \( T_{\text{eff}}\Delta(\Delta S_{\text{Ag}}^+) \)\( ^{\text{app}} \) term, Equation 2.8) would be amplified and the approximate Equation 2.10 and so the ‘internal consistency’ of alkylbenzenes in the relative affinity ladder would not validate. The high-energy (4.7 keV, laboratory frame) He-CID measurements of the decompositions of twenty-seven Ag\(^+\) bound heterodimers of alkylbenzenes were conducted by Dr. K. M. Ng [Ng, 1999] in the MAT95S B-E sector tandem mass spectrometer. The \( \ln([\text{Ag} + \text{RBz}_1]^+ / [\text{Ag} + \text{RBz}_2]^+) \) values, presented in a relative Ag\(^+\) affinity ladder (obtained at \( T_{\text{eff}} = 1,460 \) K), is shown in Figure A4.1 in Appendix I. As shown in the Figure A4.1, the relative Ag\(^+\) affinities of the fifteen alkylbenzenes are in the same order of those obtained from low-energy Ar-CID (\( T_{\text{eff}} \)
= 1,241 K) conditions. The ‘internal consistency’ of the alkylbenzenes, obtained from both high-energy (T_{eff} = 1,460 K) and low-energy (T_{eff} = 1,241 K) CID conditions, confirms that the entropic term, T_{eff}\Delta(\Delta S)^{app}, is indeed negligibly small under different collisional excitation (T_{eff}) conditions. As will be described later, the average of the quantitative data obtained under both low-energy and high-energy conditions will be taken as the final result.

**Absolute Ag (I) Cation Affinities of Alkylbenzenes.** To determine the absolute Ag\(^+\) affinities of the alkylbenzenes (i.e. \(\Delta H_{\text{Ag}^+ + \text{RBj}^+}\)), the 1/RT_{eff} term (and so T_{eff}) has to be determined according to Equation 2.10 by a series of ‘reference’ ligands with known Ag\(^+\) affinities. However, only the Ag\(^+\) affinities of benzene (156 ± 7 kJ mol\(^{-1}\) [Chen and Armentrout, 1993], 162 ± 19 kJ mol\(^{-1}\) [Ho et al., 1997]) and toluene (≤ 188 kJ mol\(^{-1}\) [Willey et al., 1991]) have been reported in literature. On the other hand, the theoretical *ab initio* Ag\(^+\) affinity value of benzene (156.7 kJ mol\(^{-1}\)), determined at the approximate CCSD(T)/[HW(f),6-31+G(d)] level using MP2/[HW, 3-21G(d)] optimized geometry by our research group, was found to be in good agreement with the experimental values from Chen and Armentrout [Chen and Armentrout, 1993] and Ho et al. [Ho et al., 1997]. This suggests that the theoretical protocol adopted by our research group might be able to reproduce experimental Ag\(^+\) affinities of alkylbenzenes with a reasonable degree of accuracy. Hence, we attempted to use theoretical Ag\(^+\) affinities of selected alkylbenzenes as ‘reference’ values to establish the experimental absolute affinities of other alkylbenzenes. The theoretical Ag\(^+\) affinities of Bz, Tol, EtBz, \(m\)-Xy, 1,3,5-Me\(_3\)Bz, 1,2,4,5-Me\(_4\)Bz, Me\(_2\)Bz and Me\(_6\)Bz were listed in Table 4.1 (and Figure 4.3 for the optimized geometries).
Figure 4.3  Optimized geometries of Ag⁺-RBz complexes determined at the MP2/[HW, 3-21G(d)] level of theory. $r_0$ represents the distance (in Å) from Ag⁺ to the centre of the aromatic ring that lies in the plane of the carbon atoms, and $\theta$ represents the angle of deviation (in degrees) of Ag⁺ to the vertical line perpendicular to the center of the ring. Only selected bond lengths and angles are presented.

Thus, a 'calibration' plot of low-energy (eV, laboratory frame / scale) Ar-CID cumulative $ln([Ag + RBz]^+ / [Ag + Bz]^+)$ values (Figure 4.2) versus the theoretically calculated affinities ($\Delta H_{[Ag + RBz]^+}$) of Tol, $m$-Xy, 1,3,5-Me₃Bz and Me₅Bz (Table 4.1) was constructed, and the Ag⁺ affinities of other alkylbenzenes were determined by interpolation and extrapolation of the plot (Figure 4.4). Similarly, another 'calibration' plot based on $ln([Ag + RBz]^+ / [Ag + Bz]^+)$ obtained under high-energy (4.7 keV, laboratory frame / scale) He-CID conditions (Figure A4.1, Appendix I) is
shown in Figure A4.2 (Appendix I). From the slopes and y-intercepts of the plots, the effective temperatures were found to be 1,241 and 1,460 K, respectively. The absolute Ag⁺ affinities (in kJ mol⁻¹) at 0 K of alkylbenzenes, together with the experimental uncertainties determined by linear regression analysis, are summarized in Table 4.1.

![Graph showing cumulative ln([Ag + RBz⁺] / [Ag + Bz⁺]) values under low-energy Ar-CID conditions versus theoretical Ag⁺ affinities of Tol, m-Xy, 1,3,5-Me₃Bz, Me₅Bz, and other alkylbenzenes.](image)

**Figure 4.4** Calibration plot of cumulative \( \ln([\text{Ag} + \text{RBz}^+] / [\text{Ag} + \text{Bz}^+]) \) values obtained under low-energy Ar-CID (Figure 4.2, \( T_{\text{eff}} = 1,241 \) K) conditions versus theoretical Ag⁺ affinities of Tol, m-Xy, 1,3,5-Me₃Bz, Me₅Bz (●) at 0 K. The absolute Ag⁺ affinities of other alkylbenzenes (○) were determined by interpolation and extrapolation of the plot.
Table 4.1  Experimental and theoretical Ag⁺ affinities, ΔH₀, at 0 K (in kJ mol⁻¹) of alkylbenzenes

<table>
<thead>
<tr>
<th>RBz</th>
<th>ΔH₀ (Expt)</th>
<th>ΔH₀ (Theory)</th>
<th>ΔH₀ (Theory) - ΔH₀ (Expt)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Low-energy Ar-CID</td>
<td>High-energy He-CID</td>
<td>Average²</td>
</tr>
<tr>
<td>Bz</td>
<td>157.2 ± 1.1 (2.2)</td>
<td>156.9 ± 0.9 (2.1)</td>
<td>157.1 ± 1.4</td>
</tr>
<tr>
<td>Tol</td>
<td>-</td>
<td>-</td>
<td>168.3d</td>
</tr>
<tr>
<td>EtBz</td>
<td>175.1 ± 0.6 (1.3)</td>
<td>176.7 ± 0.5 (1.1)</td>
<td>175.9 ± 0.8</td>
</tr>
<tr>
<td>m-Xy</td>
<td>-</td>
<td>180.6</td>
<td>180.6d</td>
</tr>
<tr>
<td>n-PrBz</td>
<td>181.1 ± 0.5 (1.1)</td>
<td>183.2 ± 0.4 (0.9)</td>
<td>182.2 ± 0.7</td>
</tr>
<tr>
<td>i-PrBz</td>
<td>181.5 ± 0.5 (1.1)</td>
<td>183.6 ± 0.4 (0.9)</td>
<td>182.3 ± 0.7</td>
</tr>
<tr>
<td>n-BuBz</td>
<td>185.3 ± 0.5 (1.0)</td>
<td>187.4 ± 0.4 (0.9)</td>
<td>186.4 ± 0.6</td>
</tr>
<tr>
<td>i-BuBz</td>
<td>185.8 ± 0.5 (1.0)</td>
<td>188.4 ± 0.4 (0.9)</td>
<td>187.5 ± 0.6</td>
</tr>
<tr>
<td>s-BuBz</td>
<td>186.8 ± 0.5 (1.0)</td>
<td>189.5 ± 0.4 (0.9)</td>
<td>188.3 ± 0.6</td>
</tr>
<tr>
<td>t-BuBz</td>
<td>188.2 ± 0.5 (1.0)</td>
<td>189.7 ± 0.4 (0.9)</td>
<td>189.0 ± 0.6</td>
</tr>
<tr>
<td>1,3,5-Me₃Bz</td>
<td>-</td>
<td>192.6d</td>
<td>192.6</td>
</tr>
<tr>
<td>n-PnBz</td>
<td>192.7 ± 0.5 (1.0)</td>
<td>194.7 ± 0.4 (1.0)</td>
<td>193.7 ± 0.7</td>
</tr>
<tr>
<td>1,2,4,5-Me₄Bz</td>
<td>199.7 ± 0.6 (1.3)</td>
<td>200.8 ± 0.5 (1.2)</td>
<td>200.3 ± 0.8</td>
</tr>
<tr>
<td>Me₅Bz</td>
<td>-</td>
<td>209.1d</td>
<td>209.1</td>
</tr>
<tr>
<td>Me₆Bz</td>
<td>216.7 ± 1.1 (2.1)</td>
<td>216.2 ± 0.8 (1.9)</td>
<td>216.5 ± 1.3</td>
</tr>
</tbody>
</table>

² Experimental Ag⁺ affinities, ΔH₀, determined from low-energy Ar-CID calibration plot (Figure 4.4, T eff = 1,241 K). The uncertainties are given as ± SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot.

² Experimental Ag⁺ affinities, ΔH₀, determined from high-energy He-CID calibration plot (Figure A4.2 in Appendix I, T eff = 1,460 K). The uncertainties are given as ± SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot.

² Average ± SD of low-energy Ar-CID and high-energy He-CID values, with standard deviation obtained from the equation s² = s₁² + s₂², where s₁² and s₂² are the SD of low-energy Ar-CID and high-energy He-CID values.

² Theoretical Ag⁺ affinities of toluene (Tol), m-xylene (m-Xy), 1,3,5-trimethylbenzene (1,3,5-Me₃Bz), pentamethylbenzene (Me₅Bz) at 0 K were adopted as reference Ag⁺ affinity values.

² Theoretical Ag⁺ affinities, ΔH₀, calculated at the CCSD(T)/[HW(0), 6-31+G(d)] level using MP2/[HW, 3-21G(d)] optimized geometries.
The experimental absolute affinities obtained under high-energy He-CID and low energy Ar-CID conditions are found to be in close agreement with each other, with a mean absolute deviation (MAD) of 1.5 kJ mol\(^{-1}\) only. In addition, our experimental Ag\(^{+}\) binding affinities of Bz, EtBz, 1,2,4,5-Me\(_4\)Bz, Me\(_6\)Bz are in excellent agreement (within ± 3 kJ mol\(^{-1}\)) with their theoretical calculated values, lending support to the accuracy of the quantitative experimental and theoretical data obtained in this study.

4.2.2 Experimental Validation of Absolute Ag (I) Cation Affinities of Alkylbenzenes

The experimental Ag\(^{+}\) affinities of acetone (160 ± 19 kJ mol\(^{-1}\) [Ho et al., 1997] and 159 ± 6 kJ mol\(^{-1}\) [El Aribi et al., 2002a]) and acetonitrile (ACN) (162 kJ mol\(^{-1}\) [Shoeib et al., 2001] and 171 ± 8 kJ mol\(^{-1}\) [El Aribi et al., 2002b], determined recently either by the radiative association kinetic analysis or threshold-CID method, were found to be comparable to those of alkylbenzenes. Hence, we attempted to re-measure the Ag\(^{+}\) affinities of these two compounds (ligands) by the independent extended kinetic method measurements using our experimental Ag\(^{+}\) affinities of alkylbenzenes as reference values. By seeking agreement between our extended kinetic method values with these literature values, the Ag\(^{+}\) affinities of alkylbenzenes found in this study can then be validated.

In the extended kinetic method measurements, the Ag\(^{+}\) bound heterodimers of acetone/ACN (L\(_n\)) and reference alkylbenzenes (RBz, including Tol, EtBz, m-Xy, i-PrBz and t-BuBz), [L\(_n\) + Ag + RBz]\(^{+}\), were allowed to dissociate under different collisional excitation energies in the collision cell of a triple quadrupole mass
A typical low-energy CID (15eV, laboratory frame) MS/MS mass spectrum of [ACN + Ag + Tol]\(^+\) heterodimer is shown in Figure 4.5.

![Graph showing mass/charge ratios and ion abundances](image)

**Figure 4.5** Triple quadrupole MS/MS (15eV, laboratory frame) spectrum of the [ACN + Ag + Tol]\(^+\) heterodimer using argon as collision gas (ACN = acetonitrile and Tol = toluene).

According to the protocol of the extended kinetic method, a first plot of \(\ln([Ag + ACN]^+ / [Ag + RBz]^+)\) values versus \([\Delta H_{[Ag + RBz]^+} - \Delta H_{AVG}]\) yields a regression line with a slope of \(-1/RT_{eff}\) and an intercept of \([\Delta G_{[Ag + ACN]^+ - \Delta H_{AVG}] / RT_{eff}\) (Figure 4.6a). A second plot was obtained by plotting \([\Delta G_{[Ag + ACN]^+ - \Delta H_{AVG}] / RT_{eff}\) against \(1/RT_{eff}\) (Figure 4.6b), and the Ag\(^+\) affinity of ACN, \(\Delta H_{[Ag + ACN]^+},\) and the apparent entropic change term, \(\Delta(S_{Ag^+})_{app}\), can then be found (Equation 2.11). The Ag\(^+\) affinity of acetone was also determined in a similar way (Figure A4.3, Appendix I), and the results are summarized in Table 4.2, together with the Ag\(^+\) affinity of benzene obtained by the standard kinetic method measurements.
Figure 4.6  (a) Plot of $\ln([\text{Ag} + \text{ACN}]^+ / [\text{Ag} + \text{RBz}]^+) \times [\Delta H_{[\text{Ag} + \text{RBz}]^+} - \Delta H_{\text{Avg}}]$ at different collision energies and (b) plot of $[\Delta G_{\text{app}}^{\text{Ag} + \text{ACN}]^+} - \Delta H_{\text{Avg}}] / RT_{\text{eff}}$ versus $1/RT_{\text{eff}}$ for the heterodimers [ACN + Ag + RBz]$^+$ (ACN = acetonitrile and RBz = Tol, EtBz, m-Xy, i-PrBz, $t$-BuBz). The Ag$^+$ affinity of ACN, $\Delta H_{[\text{Ag} + \text{ACN}]^+}$, and the $\Delta(\Delta S_{\text{Ag}^+})_{\text{app}}$ term are expressed as unweighted values.
### Table 4.2 Experimental Ag⁺ affinities, ΔH₀, at 0K (kJmol⁻¹) of benzene, acetone and acetonitrile

<table>
<thead>
<tr>
<th>L⁻</th>
<th>This work</th>
<th>Literature values</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>ΔH₀ (Expt)</td>
<td>(ΔS_{Ag⁺})_{app}</td>
</tr>
<tr>
<td>Bz</td>
<td>157.1 ± 1.4&lt;sup&gt;a&lt;/sup&gt;</td>
<td>–</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Acetone</td>
<td>151.6 ± 1.3 (2.4)&lt;sup&gt;b&lt;/sup&gt;</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>ACN</td>
<td>162.4 ± 1.8 (3.2)&lt;sup&gt;b&lt;/sup&gt;</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<sup>a</sup> Experimental Ag⁺ affinity, ΔH₀, determined by the standard kinetic method as shown in Table 4.1.

<sup>b</sup> Experimental Ag⁺ affinities, ΔH₀, determined by the extended kinetic method (Figures 4.6 and A4.3).

<sup>c</sup> Radiative associative kinetic measurements, ΔH₀ [Ho et al., 1997].

<sup>d</sup> Threshold-CID, ΔH₀ [Cheng and Armentrout, 1993].

<sup>e</sup> Threshold-CID, ΔH₀ [El Aribi et al., 2002a].

<sup>f</sup> Threshold-CID, ΔH₀ [Shoeib et al., 2001].

<sup>g</sup> Threshold-CID, ΔH₀ [El Aribi et al., 2002b].

<sup>h</sup> <i>ab initio</i> calculation at MP2 level, ΔH₀ [Ho et al., 1997].

<sup>i</sup> [Baushlicher, 1992].

<sup>j</sup> DFT calculation at B3LYP/DZVP, ΔH₀ [El Aribi et al., 2002a].

<sup>k</sup> DFT calculation at B3LYP/DZVP, ΔH₀ [Shoeib et al., 2001].

<sup>l</sup> DFT calculation at B3LYP/DZVP, ΔH₀ [El Aribi et al., 2002b].

As shown in Table 4.2, our experimental values for acetone and ACN are in very good agreement (within ± 8 – 9 kJ mol⁻¹) with those obtained by the radiative association kinetic method [Ho et al., 1997] and the threshold-CID method [Shoeib et al., 2001; El Aribi et al., 2002a; 2002b]. Furthermore, our standard kinetic method Ag⁺ affinity value of benzene (157.1 kJ mol⁻¹) is also in very good agreement (within ± 5 kJ mol⁻¹) with the reported experimental values found in literature. The maximum experimental uncertainty (90% confidence interval) obtained from the standard kinetic method measurements (Table 4.1) is 2.2 kJ mol⁻¹. Combining this error with the maximum deviation of 8 – 9 kJ mol⁻¹ observed between our experimental values and the literature
values, we tend to assign an overall uncertainty of \( \pm 10 - 11 \text{ kJ mol}^{-1} \) (depending on the 'rounding off' decimal places) to the experimental absolute \( \text{Ag}^+ \) affinities of alkylbenzenes determined in this study.

4.2.3 \( \text{Ag (I) Cation-} \pi \text{ Interaction: Ag (I) Cation Bound Complexes of Alkylbenzenes} \)

Theoretical calculations on \( \text{Ag}^+\text{-Rbz} \) systems were carried out by Miss H. M. Lee of our research group. The theoretical \( \text{Ag}^+ \) affinities of alkylbenzenes, determined at the approximate \( \text{CCSD(T)}/[\text{HW(f),6-31+G(d)}] \) level of theory, are shown in Table 4.1 and the optimized geometries of selected alkylbenzenes are shown in Figure 4.3.

As illustrated in Figure 4.3, \( \text{Ag}^+ \) binds to the aromatic-\( \pi \) face of alkylbenzenes. The natural population analysis (NPA) reveals that \( \text{Ag}^+ \) retains most of its positive charge (~ 0.91) upon complexation with alkylbenzenes. The results indicate that the binding of \( \text{Ag}^+ \) to alkylbenzenes is mainly electrostatic in nature.

A plot of \( \text{Ag}^+ \) affinities of alkylbenzenes versus the total number of carbon atoms in the alkyl substituents (Figure 4.7) reveals three effects of the alkyl / methyl groups on the strength of cation-\( \pi \) interactions. Firstly, cation-\( \pi \) interaction clearly increases with increasing number of methyl substituents (Figure 4.7, solid line). Secondly, the strength of interaction increases with increasing alkyl chain length (Figure 4.7, dotted line), even though the enhancement is relatively moderate when compared to the first effect. Lastly, a minor increase in cation-\( \pi \) binding affinity with increasing
Figure 4.7  Plot of Ag⁺ affinities of alkylbenzenes versus the total number of carbons in alkyl substituents.

As cation-π binding is mainly electrostatic in nature, the influence of alkyl substituents on Ag⁺ cation-π binding affinities should be best explained in terms of electrostatic interactions such as ion-quadrupole, ion-induced dipole, and ion-dipole interactions.
Table 4.3  Quadrupole moments \((Q_{zz} \times 10^{-40} \text{ C m}^2)\), molecular polarizabilities \((\alpha, \times 10^{-24} \text{ cm}^3)\), and molecular dipole moments \((\mu, \text{D})\) of alkylbenzenes

<table>
<thead>
<tr>
<th>RBz</th>
<th>(Q_{zz}^a)</th>
<th>(\alpha^b)</th>
<th>(\mu^c)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bz</td>
<td>-24.3 (-33.3 ± 2.1)</td>
<td>8.2 (10.0)</td>
<td>0.00 (0)</td>
</tr>
<tr>
<td>Tol</td>
<td>-22.8</td>
<td>10.2 (11.8)</td>
<td>0.32 (0.375)</td>
</tr>
<tr>
<td>EtBz</td>
<td>-21.8</td>
<td>11.8 (14.2)</td>
<td>0.28 (0.59)</td>
</tr>
<tr>
<td>m-Xy</td>
<td>-22.0</td>
<td>12.0 (14.2)</td>
<td>0.30 (0.33)</td>
</tr>
<tr>
<td>n-PrBz</td>
<td>–</td>
<td>–</td>
<td>(0)</td>
</tr>
<tr>
<td>i-PrBz</td>
<td>–</td>
<td>(16.0)</td>
<td>(0.79)</td>
</tr>
<tr>
<td>n-BuBz</td>
<td>–</td>
<td>–</td>
<td>(0)</td>
</tr>
<tr>
<td>i-BuBz</td>
<td>–</td>
<td>–</td>
<td>(0.31)</td>
</tr>
<tr>
<td>s-BuBz</td>
<td>–</td>
<td>–</td>
<td>(0)</td>
</tr>
<tr>
<td>t-BuBz</td>
<td>–</td>
<td>(17.8)</td>
<td>(0.83)</td>
</tr>
<tr>
<td>1,3,5-Me(_3)Bz</td>
<td>-21.3 (-32.1 ± 2.7)</td>
<td>13.8 (15.5)</td>
<td>0.00 (0)</td>
</tr>
<tr>
<td>n-PnBz</td>
<td>–</td>
<td>–</td>
<td>–</td>
</tr>
<tr>
<td>1,2,4,5-Me(_4)Bz</td>
<td>-20.0</td>
<td>15.6 (17.3)</td>
<td>0.00</td>
</tr>
<tr>
<td>Me(_2)Bz</td>
<td>-18.5</td>
<td>17.3 (19.1)</td>
<td>0.36</td>
</tr>
<tr>
<td>Me(_3)Bz</td>
<td>-17.1 (-24.1 ± 2.8)</td>
<td>19.1 (20.9)</td>
<td>0.00</td>
</tr>
</tbody>
</table>

\(^a\) Theoretical quadrupole moment, \(Q_{zz}\), calculated at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level of theory, in units of \(10^{-40} \text{ C m}^2\). Experimental quadrupole moments were determined by induced-birefringence method (measured in tetrachloromethane (CCl\(_4\)) at 298 K) [Vrbancici and Ritchie, 1980].

\(^b\) Theoretical molecular polarizability, \(\alpha\), calculated at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level of theory, in units of \(10^{-24} \text{ cm}^3\). Experimental molecular polarizabilities taken from [Lide, 2002-2003] are shown in the parenthesis.

\(^c\) Theoretical molecular dipole moment, \(\mu\), calculated at the MP2/3-21G(d) level of theory, in units of debye D. Experimental molecular dipole moments taken from [Lide, 2002-2003] are shown in the parenthesis.
Ion-Quadrupole Interaction. The occurrence of cation-π interaction has been attributed mainly to the attractive interaction between a cation and the relative large quadrupole moment of the benzene π-ring (refer to Section 1.2 and Footnote 4.1). However, experimental quadrupole moments are only reported for a very limited number of molecules. The available experimental quadrupole moment (Q_{zz}) of alkylbenzenes, as well as the theoretical values calculated by Miss H. M. Lee at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level of theory, are summarized in Table 4.3. The alkylbenzenes are found to have large (negative) permanent quadrupole moments (ranging from -24.3 to -17.7 × 10^{-40} C m^2) that result from the delocalized π-electron density above and below the plane of the aromatic ring. It is evident that the electrostatic Ag^+ cation-π binding in Ag^+-RBz complexes arises from a large contribution of attractive ion-quadrupole interaction.

Footnote 4.1 Molecular quadrupole moment is a measure of the non-uniform distribution of charge within a molecule, relative to a particular molecular axis [Williams, 1993]. A quadrupole moment of an aromatic ring can be considered as two dipole moments aligned in a 'tail-to-tail' fashion, or in the case of benzene, as assembly of π electrons or negative charges on both sides of the positive charged C-C ring centre, resulting in a negative quadrupole moment (Q_{zz} in a perpendicular (zz) axis to the plane of the aromatic ring).

Schematic diagram (edge-on view) of molecular quadrupole moment of benzene taken from [Dougherty, 1996]
Figure 4.8  Plot of $\text{Ag}^+$ affinities versus the negative values of theoretical quadrupole moment ($Q_{zz}$) of alkylbenzenes.

The theoretical quadrupole moments are in agreement with the trends of experimental values for Bz, 1,3,5-Me$_3$Bz and Me$_6$Bz (refer to references in footnotes of Table 4.3), even though the calculated values show deviations from the experimental absolute values. However, a plot of $\text{Ag}^+$ affinities versus the negative values of theoretical calculated quadrupole moment of alkylbenzenes (Figure 4.8) reveals that the $\text{Ag}^+$ binding affinity actually is increasing along with decreasing (less negative value) molecular quadrupole moments of the alkylbenzenes. This decreasing trend of quadrupole moment is counterintuitive if one applies the concept of ‘electron-donating’ characteristics of alkyl groups [Huang and Rodgers, 2002]. As the quadrupole moment measures the distribution of charge in a molecule, the decrease of quadrupole moment on successive methyl substitution can be explained in terms of the ‘group electronegativity’ arguments. The difference in
electronegativity between the ring carbon and the methyl groups in methyl-substituted benzenes is smaller to that between the ring carbon and hydrogen in benzene. Therefore, the charge distribution is more uniform in methyl-substituted benzenes than benzene itself. In other words, the quadrupole moments of methyl-substituted benzenes would be closer to zero, i.e., less negative when compared to benzene. Thus, the ion-quadrupole interaction would also be expected to decrease with successive methyl substitution, and this bears no direct relationship with the ‘electron-donating’ properties of the methyl or other alkyl substituents. Hence, the increasing trend of Ag⁺ cation-π affinities in alkylbenzenes is due to other types of electrostatic interaction other than the ion-quadrupole interaction.

**Ion-Induced Dipole Interaction.** Recent theoretical studies by Dargel et al. suggest that the electrostatic interaction in the Ag⁺-benzene complex is mostly due to charge-induced dipole interactions [Dargel et al., 1999]. This is confirmed by the work of Kim et al., who showed that both electrostatic (mainly ion-quadrupole interaction) and induction (polarization) interactions (mainly ion-induced dipole interaction) are important in contributing to the Ag⁺ cation-π binding of benzene, in the relative ratio of 4:7 [Kim et al., 2003]. Classically, a key parameter determining the magnitude of ion-induced dipole interaction is the molecular polarizability of the ligand. The observed increasing trends in methyl / alkyl substituted benzenes shown in Figure 4.7 is reproduced by a plot of Ag⁺ affinities versus the experimental molecular polarizabilities of methyl / alkyl substituted benzenes as shown in Figure 4.9. This clearly illustrates that the increasing trend of Ag⁺ affinities of the alkylbenzenes is due to the increasing molecular polarizabilities. As the number of methyl substitution or the number of carbon atom in the alkyl substituent increases, the molecular polarizabilities of methyl / alkyl substituted benzenes also increase.
Hence, the ion-induced dipole interaction would become much stronger, with associated increase in the Ag\(^+\) cation-\(\pi\) binding affinity.

![Graph showing Ag\(^+\) affinities versus experimental molecular polarizabilities (\(\alpha\)) of alkylbenzenes.](image)

**Figure 4.9** Plot of Ag\(^+\) affinities versus experimental molecular polarizabilities (\(\alpha\)) of alkylbenzenes.

**Intramolecular Electronic Repulsion.** As shown in Figure 4.7, the observed increasing trend of Ag\(^+\) affinities of substituted benzenes levels off slowly with longer alkyl chain or increasing number of methyl substituents. One of the contributing factors would be the decreasing trend of ion-quadrupole interaction mentioned in the previous section. Moreover, for Ag\(^+\) ion has fully-filled 4d orbitals, it is likely that electronic repulsion exists between the Ag\(^+\) 4d\(^{10}\) electrons and electrons of the ligand and hence reduce the increasing trend of Ag\(^+\) affinities. The leveling off trend is also observed in Ag\(^+\) binding to polyaromatic hydrocarbons (PAH) and related systems [Ng et al., 1998; Kim et al., 1994], and is attributed to the increase in intramolecular electronic repulsion between Ag\(^+\) ion and the ligands.
**Charge-Transfer (Covalent) Interaction.** The theoretical study by Dargel et al. also showed that charge-transfer or covalent 4d^{10} (Ag^+) → π* (benzene) and π (benzene) → 5s (Ag^+) interactions are also present [Dargel et al., 1999]. Natural bond orbital (NBO) analysis also reveals that the charge-transfer interaction is small in comparison with electrostatic interaction, with π → M^+ (σ-donor interaction) exceeding considerably the M^+ → π* (4d^{10} back donation) [Dargel et al., 1999]. The relatively small covalent contribution is attributed to the large energy gap between the donor and acceptor orbitals, as reflected in the large difference in ionization energies of Ag (7.58 eV) and benzene (9.25 eV).

On the other hand, methyl or alkyl substituted benzenes with greater molecular polarizabilities (electron polarization) than benzene may facilitate the overlapping of atomic / molecular orbitals, and thus enhance covalent interaction with the Ag^+ ion. In this regard, it is worth noting that the Ag^+ affinities of the isomeric pairs of m-Xy / EtBz (181 / 176 kJ mol^{-1}), 1,3,5-Me_3Bz / i-PrBz (193 / 183 kJ mol^{-1}) and 1,2,4,5-Me_4Bz / t-BuBz (200 / 189 kJ mol^{-1}) with comparable molecular polarizabilities are quite different (Tables 4.1 and 4.3). The Ag^+ binding affinities of successive methyl-substituted benzenes (Figures 4.7 and 4.9, solid line) are higher than that of isomeric mono-alkyl substituted benzenes (Figure 4.7 and 4.9, dotted line), and the extent of differences increases with the number of methyl substitution (Figure 4.9). The difference in Ag^+ affinities of these isomeric pairs suggest that an additional effect, apart from those discussed above, exists in affecting the strength of Ag^+ cation-π interaction. Natural population analysis showed that the positive charge (~0.91) of the Ag^+ ion in the Ag^+ -RBz complexes is less than unity, indicating that charge-transfer interaction in these complexes is non-negligible. It is likely that the positive inductive effect of the methyl / alkyl substituents (a form of electron
polarization through \( \sigma \) - and \( \pi \)-bonds) [Isaacs, 1995] stabilizes the partial positive charge of the ring carbons. The inductive effect increases with the number of methyl substituents in the aromatic ring, and is additive for successive methyl substitution. The resulting increasing inductive electron donating effect could strengthen the cation-\( \pi \) interaction by dispersing the positive charge delocalized in the aromatic ring. This is a form of charge-transfer (covalent) interaction that contributes towards enhanced \( \text{Ag}^+ \) cation-\( \pi \) interaction in methyl / alkyl substituted benzenes.

**Ion-Dipole Interaction.** According to classical electrostatics, both the magnitude and the direction of the molecular dipole moment are important for strong ion-dipole interaction. Benzene is a highly symmetric molecule with zero dipole moment. With alkyl substitution, a small molecular dipole moment may be observed due to the loss of symmetry. The experimental and theoretical molecular dipole moments (\( \mu \)) of alkylbenzenes are listed in Table 4.3. The variation of the experimental molecular dipole moment is negligible small, ranging from 0 to 0.83 Debye, which is due to the small difference of electronegativities of carbon (2.55) and hydrogen (2.20) [Allred, 1961]. Besides, the molecular dipole moments of all alkylbenzenes lie essentially along the plane of the aromatic ring, whilst the \( \text{Ag}^+ \) ion anchors on the top of the aromatic-\( \pi \) face and perpendicular to the plane of the ring. Hence, an effective interaction of \( \text{Ag}^+ \) with the molecular dipole moment is not likely. This indicates that the role of the ion-dipole interaction should be minimal in the \( \text{Ag}^+ \) cation-\( \pi \) binding of alkylbenzenes.
4.2.4 Comparison of Ag (I) Cation Affinities and Proton Affinities of Alkylbenzenes

Protonation is one of the most common and extensively studied processes in chemistry. In the case of protonated RBz, the proton is covalently bonded to the aromatic carbon(s) [Olah et al., 1970; 1972; Kuck, 1990]. The bonding arises from charge-transfer interaction, with the charge on the proton ‘delocalized’ onto the aromatic ring. The large degree of charge-transfer taken place is due to the fact that proton is a bare nucleus, with a very stable but unfilled 1s orbital, which could easily take up electrons from its neighboring environment. However, in Ag⁺-RBz complexes, the Ag⁺ ion binds to the aromatic-π face of alkylbenzenes, and the interaction is mainly electrostatic in nature with minor contribution from charge-transfer (covalent) interaction. The basic difference in the nature of interaction is reflected in the experimentally measured proton affinity of Bz (750.4 kJ mol⁻¹ [Hunter and Lias, 1998]), which is nearly 5 times greater than the Ag⁺ affinity of Bz (157.1 kJ mol⁻¹, Table 4.1 and Figure 4.10).

Alkyl substitution to the benzene ring leads to an increase in both proton and Ag⁺ affinities of alkylbenzenes, but to a greater extent for Ag⁺ affinities. Using benzene as reference, the increases in Ag⁺ affinities of n-BuBz and Me₆Bz are about 19% and 38%, respectively, while those in proton affinities are only 6% and 15%, respectively (Table A4.1, Appendix I). This observation shows that on a relative basis, alkyl substituents exerted more pronounced effects on electrostatic-based Ag⁺ affinities than covalent-based proton affinities of aromatic systems. The increase in Ag⁺ binding strength is derived mainly from ion-induced dipole interaction between the Ag⁺ ion and the aromatic ring, whereas the enhanced proton binding strength is
derived from charge-transfer interaction due to delocalization of the proton positive charge at the aromatic ring.

![Plot of Ag⁺ affinities versus proton affinities of alkylbenzenes.](image)

**Figure 4.10**  Plot of Ag⁺ affinities versus proton affinities of alkylbenzenes.

Correlation of the Ag⁺ binding energies of RBz to their proton affinities is also informative. Figure 4.10 shows a plot of Ag⁺ affinities of RBz versus proton affinities with two distinct increasing trends. The Ag⁺ affinities increase with proton affinities of alkylbenzenes for both successive methyl substitution (solid line) and increasing length of alkyl substitution (dotted line), but to a greater extent for the alkyl substituents. The result suggests that the greater molecular polarizability of the methyl / alkyl substituted benzenes (Table 4.3 and Figure 4.7) could enhance the strength of Ag⁺ binding via ion-induced dipole interaction, but the effect is smaller on charge-transfer interaction. This provides further evidence that the increasing
trend of Ag⁺ affinities among alkylbenzenes is mainly due to ion-induced dipole interaction. In fact, the enhancement in Ag⁺ affinities of unbranched RBz is larger if only molecular polarizability factor is at work. This is attributed to the possibility of longer and more polarizable alkyl chain "wrapping" around the Ag⁺ ion, as is in the case of Li⁺-(n-BuOH) [Rodgers and Armentrout, 1997; Siu et al., 1998]. This ‘wrapping’ effect is more noticeable for n-butyl and n-pentylbenzenes as 4 or 5 carbon chains are favored to form a ‘cyclized’ and stable structure of Ag⁺-RBz complexes. However, this effect should be minimal in the case of protonated RBz.

4.2.5 Comparison of Ag (I) Cation Affinities with Other Metal Cation Affinities of Alkylbenzenes

Recently, many studies have been carried out to examine the origin of cation-π interactions with different metal cations. The reported experimental binding affinities of Bz and Tol for different metal cations, including the Group (I) alkali metal (Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺) and Group (II) transition metal cations (Cu⁺, Ag⁺, Au⁺), at 0 K are summarized in Table 4.4.
Table 4.4  Experimental binding affinities of Group (I) alkali metal and Group (XI) transition metal cations to benzene and toluene at 0 K ($\Delta H_0$ in kJ mol$^{-1}$)

<table>
<thead>
<tr>
<th></th>
<th>Group (I) Alkali Metal Cation</th>
<th>Group (XI) Transition Metal Cation</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Li$^+$</td>
<td>Na$^+$</td>
</tr>
<tr>
<td>Bz</td>
<td>$151.9 \pm 8.0^a$</td>
<td>$115.8 \pm 6.3^b$</td>
</tr>
<tr>
<td></td>
<td>$161.1 \pm 13.5^d$</td>
<td>$92.6 \pm 5.8^d$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$68.5 \pm 3.8^d$</td>
</tr>
<tr>
<td></td>
<td>$88.3 \pm 4.3^e$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Cu$^+$</td>
<td>$218 \pm 10^f$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$212^g$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$\leq 191^h$</td>
</tr>
<tr>
<td>Tol</td>
<td>$183.1 \pm 16.0^a$</td>
<td>$112.3 \pm 3.5^a$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>$71.3 \pm 4.2^a$</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

$^a$ Ion cyclotron resonance mass spectrometry (ICR-MS), [Woodin and Beauchamp, 1978].
$^b$ High-pressure mass spectrometry (HPMS), [Guo et al., 1990].
$^c$ HPMS, [Sunner et al., 1981].
$^d$ Threshold-CID, $\Delta H_0$, [Amicangelo and Armentrout, 2000].
$^e$ Threshold-CID, $\Delta H_0$, [Armentrout and Rodgers, 2002].
$^f$ Threshold-CID, $\Delta H_0$, [Meyer et al., 1995].
$^g$ Experimental CID modeling result, [Yang and Klippenstein, 1999].
$^h$ $\Delta H_{ess}$, [Willey et al., 1991].
$^i$ Threshold-CID, $\Delta H_0$, [Chen and Armentrout, 1993].
$^j$ Radiative association kinetic analysis, [Ho et al., 1997].
$^k$ This work, standard kinetic method $\Delta H_0$ value as shown in Table 4.1.
$^l$ FTICR-MS, [Ho and Dunbar, 1999].
$^m$ DFT calculation at B3LYP level of theory, $\Delta H_0$, [Schroeder et al., 2000].
$^n$ Threshold-CID, $\Delta H_0$, [Amunugama and Rodgers, 2000a].
As discussed earlier, and similar to the case of Group (I) alkali metal cations [Amicangelo and Armentrout, 2000; 2002a; Tsuzuki et al., 2001], the binding of Ag⁺ cation-π complexes contributes mainly from ion-quadrupole and ion-induced dipole interactions. On purely electrostatic grounds, a smaller size metal ion has a greater surface charge density and a closer approach to the π system, leading to a stronger electrostatic interaction. Although the size of Ag⁺ ion (1.26 Å [Huheey et al., 1993]) is in-between Na⁺ and K⁺ (0.60, 0.95, 1.33, 1.48, 1.69 Å for Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, respectively), the strengths of Ag⁺ binding to Bz and Tol are larger than those of Na⁺, K⁺, Rb⁺, Cs⁺, and is comparable to that of Li⁺ (i.e. Cs⁺ < Rb⁺ < K⁺ < Na⁺ < Li⁺ ~ Ag⁺). This reversal in the trend of metal cation affinity versus ionic radius is likely to be due to the presence of 4d¹⁰ electrons and more covalent interaction in the case of Ag⁺, but are insignificant in the larger alkali metal cations like Na⁺ and K⁺.

Another informative feature shown in Table 4.4 is that the Group (XI) transition metal Cu⁺ (ionic radius 0.96 Å) shows a greater affinity of benzene than Ag⁺ (ionic radius 1.26 Å), suggesting that the bindings in Cu⁺-Bz and Ag⁺-Bz complexes are indeed mainly electrostatic in origin. In contrast, Au⁺, with a larger ionic radius (1.37 Å), displays much stronger affinity towards benzene than Ag⁺ and Cu⁺. This is due to a relatively large covalent contribution in the Au⁺-Bz interaction, which has been interpreted in terms of relativistic effects [Hertwig et al., 1996; Dargel et al., 1999]. Relativity causes a significant stabilization of the Au⁺ (6s⁰5d¹⁰) valence 6s orbital, while the occupied 5d orbitals are destabilized and expanded. This effect leads to a considerable increase in ionization energy of Au (9.23 eV) as compared to Cu (7.73 eV) and Ag (7.58 eV), which facilitates more significant electron transfer from Bz towards Au⁺ (π → 6s orbitals) (for Bz has ionization energy of 9.25 eV comparable to Au). Likewise, the ability of electrons back donated from Au⁺ to Bz
(5d → π* orbitals) is also enhanced. Thus, the increase in covalent character of the Au⁺-Bz complex results in a stronger binding strength than those of Cu⁺/Ag⁺-Bz complexes (i.e. Ag⁺ < Cu⁺ < Au⁺), which are also observed in Cu⁺ / Ag⁺ / Au⁺ binding to 2-propanol, acetone and propene [Huang et al., 2001].

Schroeter et al. reported the Cr⁺, Fe⁺ and Co⁺ cation affinities of alkyl / methyl substituted benzenes by the kinetic method [Schroeter et al., 1998]. The affinity order is found to be: Co⁺ > Fe⁺ > Cr⁺ ∼ Ag⁺. This result may reflect the smaller ionic radius (hence greater ion-induced dipole interaction) of these first-row transitional metals, as well as more significant covalent contribution for Co⁺, Fe⁺ and Cr⁺ (arising from partially filled d-shell) than that of Ag⁺.

The correlation plots of available metal cation affinities of RBz (M⁺ = Li⁺, Na⁺, Cr⁺, Fe⁺, Co⁺, Ag⁺) versus experimental molecular polarizabilities (Table 4.3) are shown in Figure 4.11a and 4.11b. The nearly linear correlations (r² ~ 0.9) clearly illustrate a prevailing notion that the molecular polarizability exhibits a pronounced effect on the interaction of various metal cations binding to π systems [Ma and Dougherty; 1997; Dargel et al., 1999; Tsuzuki et al., 2001; Kim et al., 2003], particularly in Ag⁺-RBz complexes as discussed in Section 4.2.3. Hence, it is not surprising that the Ag⁺ affinities of RBz can be linearly correlated with the alkali / transition metal cation (Li⁺ / Na⁺ / Cr⁺ / Fe⁺ / Co⁺) affinities as shown in Figure 4.11c.

As shown in the plot of metal cation affinity versus molecular polarizabilities (Figure 4.11b), Ag⁺ shows a steeper slope than those from Cr⁺, Fe⁺ and Co⁺, and a 'cross-over' of the Ag⁺ and Cr⁺ lines are observed. Apart from the consideration of polarization (induction) interaction, we note the finding by Kim et al. that a comparatively large contribution of dispersion interaction (ion-induced-ion-induced
dipole interaction) is found in the Ag⁺-Bz complexes [Kim et al., 2003]. Dispersion energy is expected to increase with the polarizabilities of the metal cation as well as the ligand. The presence of 4d¹⁰ electrons of Ag⁺ is expected to increase the 'ionic polarizability' of Ag⁺ relative to that of Co⁺, Fe⁺ and Cr⁺ (partially filled d shell). Hence, the inclusion of dispersion energy effect may explain the more steep variation of Ag⁺ affinities among alkyl / methyl substituted benzenes when compared to those of Co⁺, Fe⁺ and Cr⁺ (Figure 4.11b).
Figure 4.11  Plots of metal cation (M⁺) affinities versus molecular polarizabilities of (a) alkyl substituted benzenes and (b) methyl substituted benzenes. (c) Plots of Ag⁺ affinities versus M⁺ affinities of substituted benzenes (M⁺ = Na⁺ (○), Li⁺ (●), Cr⁺ (▲), Fe⁺ (▲), Co⁺ (□), Ag⁺ (×)). The molecular polarizabilities increase in the order of (a) Bz > Tol > EtBz > t-BuBz and (b) Bz > Tol > m-Xy > 1,3,5-Me₃Bz > 1,2,4,5-Me₄Bz > Me₂Bz > Me₆Bz. (c) The Na⁺ and Li⁺ affinities calculated at MP2/6-31G** level are taken from [Tsuzuki et al., 2001] in the order of Bz > Tol > EtBz > t-BuBz. The Cr⁺, Fe⁺ and Co⁺ affinities measured by kinetic method are taken from [Schroeter et al., 1998] in the order of Bz > Tol > EtBz > m-Xy > 1,3,5-Me₃Bz > 1,2,4,5-Me₄Bz > Me₂Bz > Me₆Bz.
Chapter 5  Ag (I) Cation Affinities of Substituted Naphthalenes

5.1 Background

Naphthalene is a good model, though not exactly the same, to mimic the behavior of tryptophan (one of the commonly occurring aromatic amino acids), as well as biological assemblies containing extended π networks [Amunugama and Rodgers, 2003a]. Previous investigations of cation-π interactions involving naphthalene have been limited. In the first theoretical study by Mecozzi et al., the Na⁺ binding affinities of a variety of aromatic ligands, including naphthalene, were examined [Mecozzi et al., 1996]. The second theoretical study involved an ab initio mapping study of the interactions of naphthalene and indole with Na⁺, Mg⁺, and Al⁺ [Dunbar, 1998]. In the third study, the conformations, energetics and solvation effects of the alkali metal cation (Li⁺, Na⁺, K⁺) binding to naphthalene were investigated [Hashimoto and Ikuta, 1999]. The first experimental work was reported by Dunbar and co-workers who examined the radiative association reactions of Si⁺, Fe⁺, Cr⁺, and Mn⁺ with naphthalene [Dunbar et al., 1994]. Amunugama and Rodgers recently reported the alkali metal (Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺) cation-π binding affinity of naphthalene measured by the threshold-CID method, and also the theoretical affinities calculated at the MP2(full)/6-311+G(2d,2p)//B3LYP/6-31G* level of theory [Amunugama and Rodgers, 2003a].

The present work aims to extend our study on Ag⁺ cation-π model systems to a fused aromatic system. The Ag⁺ binding affinity of naphthalene (Nap) and substituted
naphthalenes (X-Nap, where X = H, Me, Et, i-Pr, OH, OMe, OEt, CN, NH$_2$ substituents at the 2-position) were determined by the mass spectrometric kinetic method. For simplicity of presentation, the positions of the substituents are omitted in our abbreviations and text hereafter. The effects of a fused-benzene on cation-π and non-π modes of Ag$^+$ binding are discussed.

![Graphs showing mass spectra](image)

**Figure 5.1** Low-energy ((a) 10eV and (b) 25eV, laboratory frame) Ar-CID mass spectra of Ag (I) cation bound heterodimer, [Nap + Ag + (Me-Nap)]$^+$, (Nap = naphthalene and Me-Nap = 2-methylnaphthalene).
5.2 Results and Discussion

5.2.1 Determination of Ag (I) Cation Affinities of Substituted Naphthalenes

Relative Ag (I) Cation Affinities of Substituted Naphthalenes. Similar to the case of alkylbenzenes (Chapter 4), we first attempted to establish the order of Ag⁺ affinities among substituted naphthalenes by standard kinetic method measurements. The low-energy Ar-CID of ESI generated [(X-Nap₁) + Ag + (X-Nap₂)]⁺ heterodimers to its [Ag + (X-Nap₁)]⁺ and [Ag + (X-Nap₂)]⁺ monomer complexes were measured at two different collision energies at 10eV and 25eV. Typical low-energy Ar-CID mass spectra of Ag⁺ bound heterodimer, [Nap + Ag + (Me-Nap)]⁺ (where Nap = naphthalene and Me-Nap = 2-methylnaphthalene) are shown in Figure 5.1. The corresponding natural logarithm ratios, \( \ln([\text{Ag} + (\text{Me-Nap})]^+ / [\text{Ag} + \text{Nap}]^+) \), are 1.17 and 0.93, respectively, which was found to decrease with increase in collision energy (or effective temperature, \( T_{\text{eff}} \)).

Totally, the low-energy Ar-CID of Ag⁺ bound heterodimers of seven substituted naphthalenes including naphthalene (Nap), 2-methylnaphthalene (Me-Nap), 2-ethynaphthalene (Et-Nap), 2-isopropynaphthalene (i-Pr-Nap), 2-hydroxynaphthalene (HO-Nap), 2-methoxynaphthalene (MeO-Nap), 2-ethoxynaphthalene (EtO-Nap) were measured at a collision energy of 10 eV. Similar to the case of alkylbenzenes (RBz) illustrated in Chapter 4 (Figure 4.2), the natural logarithm ratios \( \ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+) \) (where \( L_1 \) and \( L_2 \) represent the substituted naphthalenes or reference alkylbenzenes) are assembled in a relative Ag (I) cation affinity ladder (at \( T_{\text{eff}} = 1,132 \text{ K} \)) as shown in Figure 5.2. 
<table>
<thead>
<tr>
<th>Ligands, L</th>
<th>(\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+))</th>
<th>(\ln([\text{Ag} + L]^+ / [\text{Ag} + \text{Tol}]^+))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Tol</td>
<td></td>
<td>0.00 ± 0.00</td>
</tr>
<tr>
<td>Nap</td>
<td></td>
<td>0.91 ± 0.00</td>
</tr>
<tr>
<td>HO-Nap</td>
<td>1.36</td>
<td>0.91 ± 0.00</td>
</tr>
<tr>
<td>m-Xy</td>
<td>2.44</td>
<td>1.09 ± 0.03</td>
</tr>
<tr>
<td>Me-Nap</td>
<td>1.13</td>
<td>1.14 ± 0.03</td>
</tr>
<tr>
<td>MeO-Nap</td>
<td>1.52</td>
<td>1.46 ± 0.03</td>
</tr>
<tr>
<td>1,3,5-Me₃Bz</td>
<td>1.00</td>
<td>2.14 ± 0.03</td>
</tr>
<tr>
<td>Et-Nap</td>
<td>0.68</td>
<td>2.44 ± 0.03</td>
</tr>
<tr>
<td>EtO-Nap</td>
<td>0.45</td>
<td>2.58 ± 0.04</td>
</tr>
<tr>
<td>i-Pr-Nap</td>
<td>0.44</td>
<td>2.74 ± 0.05</td>
</tr>
<tr>
<td></td>
<td></td>
<td>2.94 ± 0.05</td>
</tr>
<tr>
<td></td>
<td></td>
<td>3.19 ± 0.06</td>
</tr>
</tbody>
</table>

Figure 5.2  Experimentally measured \(\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)\) values for the low-energy (10 eV, laboratory frame) Ar-CID of Ag⁺ bound heterodimers of substituted naphthalenes \((T_{\text{eff}} = 1,132 \, \text{K})\). The \(\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)\) values are the logarithm of ion intensity ratios. The data presented under the heading The \(\ln([\text{Ag} + L]^+ / [\text{Ag} + \text{Tol}]^+)\) are average ± standard deviation of cumulative values expressed relative to toluene (Tol). The values in ( ) are the arithmetic difference of measured values having a common reference ligands.

The relative Ag⁺ affinities of some substituted naphthalene pairs, which differ in their molecular mass (or m/z value) by 2 units, such as EtO-Nap and i-Pr-Nap, cannot be measured directly by pairing them in a heterodimer due to overlapping of \(^{107}\text{Ag}^+\) and \(^{109}\text{Ag}^+\) isotopic peaks. Their relative Ag⁺ affinities were determined from their respective heterodimers with other substituted naphthalenes or reference alkylbenzenes. For example, \(\ln([\text{Ag} + (i-\text{Pr-Nap})]^+ / [\text{Ag} + (\text{Et-Nap})]^+)\) and \(\ln([\text{Ag} + (\text{EtO-Nap})]^+ / [\text{Ag} + (\text{Et-Nap})]^+)\) values were 0.44 and 0.20, respectively. The
arithmetic difference between these two values, corresponding to $\ln([\text{Ag} + (i-\text{Pr-Nap})]^{\dagger}/[\text{Ag} + (\text{EtO-Nap})]^{\dagger})$, was 0.24 (i.e., 0.44 - 0.20). Another way of obtaining this ratio is from the arithmetic difference between the $\ln([\text{Ag} + (i-\text{Pr-Nap})]^{\dagger}/[\text{Ag} + (\text{Me-Nap})]^{\dagger})$ and $\ln([\text{Ag} + (\text{EtO-Nap})]^{\dagger}/[\text{Ag} + (\text{Me-Nap})]^{\dagger})$ values at 1.00 and 0.73, respectively, yielding a $\ln([\text{Ag} + (i-\text{Pr-Nap})]^{\dagger}/[\text{Ag} + (\text{EtO-Nap})]^{\dagger})$ value of 0.27 (i.e. 1.00 - 0.73). Similar treatment to another pairs of heterodimer was carried out. The average $\ln([\text{Ag} + (i-\text{Pr-Nap})]^{\dagger}/[\text{Ag} + (\text{EtO-Nap})]^{\dagger})$ values was 0.25 (SD = 0.02, $n = 3$), which was adopted as the final value.

Our results show that the relative affinity ladder of substituted naphthalenes shown in Figure 5.2 is internally consistent. Taking the [Nap + Ag + (MeO-Nap)]$^{\dagger}$ heterodimer as an example, the direct measurement of $\ln([\text{Ag} + \text{Nap}]^{\dagger}/[\text{Ag} + (\text{MeO-Nap})]^{\dagger})$ value was 1.45, which is close to the summation of the $\ln([\text{Ag} + L_1]^{\dagger}/[\text{Ag} + L_2]^{\dagger})$ value of the four successive stair-steps from Nap to MeO-Nap, i.e., 0.18 + 0.37 + 0.68 + 0.30 = 1.53.

To our initial surprise, the internal consistency (i.e. $\Delta(\Delta S_{Ag^+})^{app} \approx 0$, Equation 2.8) can be extended to dissociation of heterodimers containing 2-cyanonaphthalene (CN-Nap) and 2-aminonaphthalene (NH$_2$-Nap), which bind to Ag$^+$ in the non-$\pi$ mode (refer to Section 5.2.2), different from the cation-$\pi$ binding mode of other substituted naphthalenes (i.e. Nap, Me-Nap, Et-Nap, HO-Nap, MeO-Nap, EtO-Nap) of the relative affinity ladder. This is illustrated separately for clarity of presentation in the relative affinity ladder obtained at 25 eV (Figure A5.1, Appendix II). As mentioned in Section 2.4, the approximation, $\Delta(\Delta S_{Ag^+})^{app} \approx 0$, may not hold for dissociating heterodimers comprised of bound ligands with different binding modes. Thus, this finding is likely to be coincidental, but allows the Ag$^+$ affinities of CN-Nap and
NH$_2$-Nap to be measured from the relative affinity ladder shown in Figure 5.2 and Figure A5.1.

The relative Ag$^+$ affinities of the nine substituted naphthalenes, expressed in terms of $\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ ratios, were found in the order:


Figure 5.3  Low-energy Ar-CID ((a) 10 eV, (b) 25 eV, laboratory frame) mass spectra of the [Nap + Ag + (m-Xy)]$^+$ heterodimer (Nap = naphthalene and m-Xy = m-xylene).
Absolute Ag (I) Cation Affinities of Substituted Naphthalenes. To establish the absolute Ag⁺ affinities of substituted naphthalenes, we attempted to include alkylbenzenes of known Ag⁺ affinities in the relative affinity ladder of substituted naphthalenes shown in Figure 5.2 and A5.1. The chosen ‘reference’ alkylbenzenes are toluene (Tol), m-xylene (m-Xy) and 1,3,5-trimethylbenzene (1,3,5-Me₃Bz), which show Ag⁺ affinities comparable to that of substituted naphthalenes. Typical examples of low-energy Ar-CID mass spectra of the Ag⁺ bound heterodimer between naphthalene and an alkylbenzene (Nap + Ag + (m-Xy))⁺, are shown in Figure 5.3. Their ln([Ag + L₁]⁺ / [Ag + L₂]⁺) intensity ratios are also included in the relative Ag⁺ affinity ladder shown in Figure 5.2 and A5.1 (Appendix II).

With inclusion or ‘embedding’ of reference alkylbenzenes, the relative Ag⁺ affinity ladder shown in Figure 5.2 and A5.1 were found to be internally consistent. Following the same procedure as described in Section 4.2.1, a calibration plot of cumulative ln([Ag + L]⁺ / [Ag + Tol]⁺) values obtained under low-energy (10eV, laboratory frame) Ar-CID condition (Figure 5.2) versus the Ag⁺ affinities of reference alkylbenzenes (Tol, m-Xy and 1,3,5-Me₃Bz) at 0K is shown in Figure 5.4. The high value of correlation (R² = 0.99) in Figure 5.4 is a direct result of internal consistency of the relative Ag⁺ affinity ladder shown in Figure 5.2. The T_eff values of 1,132 K and the absolute Ag⁺ affinities of X-Nap could be derived from the slope (i.e. 1/RT_eff = 0.11) and y-intercept (i.e. −ΔH_{Ag + Tol}^+/RT_eff = −17.83) according to Equation 2.10. Similar calibration plots obtained for the relative affinity ladder determined at a collision energy of 25eV (laboratory frame) as shown in Figure A5.2 (Appendix II) with T_eff values at 1,563 K and 1,454 K. The absolute Ag⁺ affinities so obtained are also summarized in Table 5.1.
Figure 5.4  Calibration plot of cumulative $ln([\text{Ag} + \text{L}]^+ / [\text{Ag} + \text{Tol}]^+)$ values (Figure 5.2) obtained under low-energy (10eV, laboratory frame) Ar-CID ($T_{\text{eff}} = 1,132$K) conditions versus Ag$^+$ affinities of reference alkylbenzenes of Tol, $m$-Xy, 1,3,5-Me$_3$Bz (●) at 0K. The absolute Ag$^+$ affinities of substituted naphthalenes (○) were determined by interpolating and extrapolating of the plot.

As shown in Table 5.1, excellent agreement is found between the Ag$^+$ affinities determined at the two collision energies of 10eV and 25eV, with deviations within ± 2 kJ mol$^{-1}$. The average affinities of X-Nap obtained at the low-energy Ar-CID measurements were adopted as our final results, and they will be used in our further discussions hereafter. The maximum uncertainty of the linear ‘calibration plots’ (Figure 5.4 and A5.2) is 2.9 kJ mol$^{-1}$ (90% C.L.). Hence, combining this error with the estimated uncertainty of the reference alkylbenzenes (10 – 11 kJ mol$^{-1}$, Section
4.2.2), the overall experimental uncertainty of the measured Ag$^+$ affinities of X-Nap is estimated to be $\pm 13 - 14 \text{ kJ mol}^{-1}$.

<table>
<thead>
<tr>
<th>Table 5.1</th>
<th>Experimental Ag$^+$ affinities, $\Delta H_0$, at 0K (in kJ mol$^{-1}$) of substituted naphthalenes</th>
</tr>
</thead>
<tbody>
<tr>
<td>X-Nap</td>
<td>Low-energy (10 eV, lab frame) Ar-CID$^a$</td>
</tr>
<tr>
<td>Nap</td>
<td>176.4 ± 0.7 (1.6)</td>
</tr>
<tr>
<td>Me-Nap</td>
<td>188.0 ± 0.9 (1.8)</td>
</tr>
<tr>
<td>Et-Nap</td>
<td>193.6 ± 1.1 (2.4)</td>
</tr>
<tr>
<td>i-Pr-Nap</td>
<td>197.9 ± 1.4 (2.9)</td>
</tr>
<tr>
<td>HO-Nap</td>
<td>178.1 ± 0.7 (1.5)</td>
</tr>
<tr>
<td>MeO-Nap</td>
<td>190.8 ± 1.0 (2.1)</td>
</tr>
<tr>
<td>EtO-Nap</td>
<td>195.5 ± 1.2 (2.6)</td>
</tr>
<tr>
<td>CN-Nap</td>
<td>–</td>
</tr>
<tr>
<td>NH$_2$-Nap</td>
<td>–</td>
</tr>
</tbody>
</table>

$^a$ Experimental Ag$^+$ affinities, $\Delta H_0$, measured by low-energy (10eV, laboratory frame) Ar-CID calibration plot (Figure 5.4, $T_{eff} = 1,132 \text{ K}$). The uncertainties are given as ± SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot.

$^b$ Experimental Ag$^+$ affinities, $\Delta H_0$, measured by low-energy (25eV, laboratory frame) Ar-CID calibration plot (Figure A5.2, in Appendix II, $T_{eff} = 1,563 \text{ K}$ (for alkyl / NH$_2$ substituted naphthalenes) or 1,454 K (for alkoxyl / CN substituted naphthalenes)). The uncertainties are given as ± SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot.

$^c$ Average ± SD of two low-energy Ar-CID values, with standard deviation obtained from the equation $s^2 = s_1^2 + s_2^2$, where $s_1^2$ and $s_2^2$ are the SD of low-energy Ar-CID 10 eV (laboratory frame) and 25 eV (laboratory frame) values.

Because of the ‘co-incidental’ inclusion of –CN and –NH$_2$ substituted naphthalenes into the relative affinity ladder in Figure A5.1 (Appendix II), we carried out confirmatory measurements on the Ag$^+$ affinities of CN-Nap and NH$_2$Nap by the extended kinetic method, with explicit measurement of the $\Delta(\Delta S_{Ag})_{N,pp}$ term and
using RBz including $m$-Xy, 1,3,5-Me$_2$Bz, 1,2,4,5-Me$_4$Bz and Me$_3$Bz as the references compounds. Typical low-energy Ar-CID mass spectra of [(CN-Nap) + Ag + (1,3,5-Me$_3$Bz)]$^+$ heterodimers are shown in Figure 5.5. The Ag$^+$ affinity of CN-Nap, as well as the $\Delta(\Delta S_{Ag^+})^{app}$ value can be obtained from the corresponding first plot and second plots of the extended kinetic method measurements are shown in Figure 5.6. The Ag$^+$ affinity of NH$_2$-Nap were determined in a similar way (Figure A5.3 and Figure A5.4, Appendix II), and the results are summarized in Table 5.2.

As listed in Table 5.2, the agreements between the experimental affinities obtained from the standard and extended kinetic methods is very good, with deviation of ±3 kJ mol$^{-1}$. Also, the average of the apparent entropic term contribution in the extended kinetic method measurements, $\Delta(\Delta S_{Ag^+})^{app}$, was found to be small (<±2 J K$^{-1}$ mol$^{-1}$), confirming the result of the standard kinetic method measurements.

Table 5.2  Experimental Ag$^+$ affinities, $\Delta H_0$, at 0K (in kJ mol$^{-1}$) of 2-cyanonaphthalene and 2-aminonaphthalene

<table>
<thead>
<tr>
<th>X-Nap</th>
<th>Standard Kinetic Method</th>
<th>Extended Kinetic Method</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>$\Delta H_0$ (Exp) $^a$</td>
<td>$\Delta H_0$ (Exp) $^b$</td>
</tr>
<tr>
<td>CN-Nap</td>
<td>194.8 ± 0.6 (1.4)</td>
<td>192.7 ± 0.2 (0.4)</td>
</tr>
<tr>
<td>NH$_2$-Nap</td>
<td>211.1 ± 0.5 (1.0)</td>
<td>214.3 ± 0.7 (1.3)</td>
</tr>
</tbody>
</table>

$^a$ Experimental affinities, $\Delta H_0$, determined by the standard kinetic method as shown in Table 5.1.

$^b$ Experimental affinities, $\Delta H_0$, determined by the extended kinetic method as shown in Figure 5.6 and A5.4 (Appendix II). Weighted results with experimental uncertainties given as ± SD (90% confidence interval) in the format of reference [Armentrout, 2000].
Figure 5.5  Triple quadrupole MS/MS spectrum of the [(CN-Nap) + Ag + (1,3,5-Me₃Bz)]⁺ heterodimer using argon as collision gas at a various collision energies (laboratory frame) (CN-Nap = 2-cyanonaphthalene and 1,3,5-Me₃Bz = 1,3,5-trimethylbenzene).
Figure 5.6  (a) Plot of $\ln([\text{Ag} + (\text{CN-Nap})^+] / [\text{Ag} + \text{RBz}]^+)$ versus $[\Delta H_{\text{Ag} + \text{RBz}}]^+ - \Delta H_{\text{Avg}}$ at different collision energies and (b) plot of $[\Delta G_{\text{app}}^{\text{pp}}]_{\text{Ag} + (\text{CN-Nap})^+} - \Delta H_{\text{Avg}}] / RT_{\text{eff}}$ versus $1/RT_{\text{eff}}$ for the heterodimers [(CN-Nap) + Ag + RBz]$^+$ (CN-Nap = 2-cyanonaphthalene and RBz = $m$-Xy, 1,3,5-Me$_3$Bz, 1,2,4,5-Me$_4$Bz, Me$_2$Bz). The Ag$^+$ affinity of CN-Nap, $\Delta H_{\text{Ag} + (\text{CN-Nap})^+}$, and the entropy changes, $\Delta (\Delta S_{\text{Ag}^+})^{\text{pp}}$, are expressed as unweighted values.
5.2.2 Comparison of Experimental and Theoretical Ag (I) Cation Affinities of Substituted Naphthalenes

To provide further understanding on the binding modes and nature of Ag\(^+\) binding, related theoretical calculations on the Ag\(^+\)-(X-Nap) binding geometries and affinities were carried out by Miss H. M. Lee. The binding affinities were calculated at the CCSD(T)/[HW(f), 6-31+G(d)] level using optimized geometries at the MP2/[HW, 3-21G(d)] level. The theoretical determined Ag\(^+\) affinities of Nap, Me-Nap, Et-Nap, t-Pr-Nap, HO-Nap, MeO-Nap, EtO-Nap, CN-Nap and NH\(_2\)-Nap at 0K are summarized in Table 5.3, and their most stable structures are shown in Figure 5.7.

As in the case of substituted benzenes, Ag\(^+\) may bind to substituted naphthalenes in the cation-\(\pi\) and non-\(\pi\) mode of binding. Furthermore, there are two \(\pi\) sites in substituted naphthalene. Here, we denote the naphthalene benzene-\(\pi\) ring with the substituent attached as 'ring I' and the unsubstituted one as 'ring II'. The theoretical results show that Ag\(^+\) preferentially binds to substituted naphthalenes in the following manner (Table 5.3):

Cation-\(\pi\), ring I: electron donating alkyl or alkoxy substituents attached

Cation-\(\pi\), ring II: electron donating \(-\text{OH}\) substituent

Non cation-\(\pi\): nitrogen containing \(-\text{CN}, -\text{NO}_2\) and \(-\text{NH}_2\) substituents

As shown in Table 5.3, our experimental order of Ag\(^+\) affinities of alkyl and alkoxy substituted naphthalenes is in agreement with the theoretical prediction, i.e. Nap < HO-Nap < Me-Nap < MeO-Nap < Et-Nap < EtO-Nap < t-Pr-Nap, even though the theoretical protocol could be slightly underestimating the absolute Ag\(^+\) affinities of alkyl, hydroxyl and alkoxy substituted naphthalenes (by an average of 3 kJ mol\(^{-1}\)).
Furthermore, the difference in Ag\(^+\) binding affinity of benzene-\(\pi\) ring I (with substituent attached) and ring II (without substituent attached) is quite small (\(\sim 5.4\) kJ mol\(^{-1}\)). Thus, under low-energy CID conditions, the kinetic method measurements are likely to yield Ag\(^+\) binding affinities of alkyl and alkoxy substituted naphthalenes in their most stable Ag\(^+\) cation-\(\pi\) binding modes (ring I), though the possible occurrence of Ag\(^+\) binding to the less stable benzene-\(\pi\) ring II in both the Ag\(^+\) bound monomer and heterodimer complexes, cannot be entirely excluded.

The experimental Ag\(^+\) affinity of 2-cyanonaphthalene (CN-Nap) and 2-aminonaphthalene (NH\(_2\)-Nap) determined by standard kinetic method are in very good agreement with the theoretical values, with very acceptable deviations at 3.3 and -7.0 kJ mol\(^{-1}\), respectively, and well within the estimated experimental uncertainty \(\pm 13 - 14\) kJ mol\(^{-1}\) of our kinetic method measurements. We note that while the experimental Ag\(^+\) affinity value of EtO-Nap (195.3 kJ mol\(^{-1}\)) is slightly larger than that of CN-Nap (194.8 kJ mol\(^{-1}\)), the theoretical values for the most stable binding modes is of a different order CN-Nap (198.1 kJ mol\(^{-1}\)) > EtO-Nap (192.7 kJ mol\(^{-1}\)). Given the differences are small, the discrepancy most likely arises from the combined errors of experimental measurements and theoretical calculations. Therefore, we tend to conclude that CN-Nap and EtO-Nap have comparable and similar Ag\(^+\) affinities. On the whole, the very good agreements between experimental and theoretical values provide confidence to the results obtained in the present study.
Figure 5.7  Optimized geometries of the most stable Ag⁺-(X-Nap) complexes at MP2/[HW, 3-21G(d)] level of theory.  \( r \) represents the distance (in Å) from Ag⁺ to the centre of the aromatic ring that lies in the plane of the carbon atoms, and \( \theta \) represents the angle of deviation (in degrees) of Ag⁺ to the vertical line perpendicular to the center of the ring. Only selected bond lengths and angles are presented.
Table 5.3 Experimental and theoretical Ag$^+$ affinities, $\Delta H_0$, at 0K (in kJ mol$^{-1}$) of substituted naphthalenes

<table>
<thead>
<tr>
<th>X-Nap</th>
<th>$\Delta H_0$ (Exp)$^a$</th>
<th>Benzene-$\pi$ ring I$^c$</th>
<th>Benzene-$\pi$ ring II$^d$</th>
<th>Non-$\pi$ Binding$^e$</th>
<th>$\Delta H_0$ (Theory) $- \Delta H_0$ (Exp)$^f$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nap</td>
<td>176.8</td>
<td>175.3</td>
<td>175.3</td>
<td>-</td>
<td>-1.5</td>
</tr>
<tr>
<td>Me-Nap</td>
<td>188.5</td>
<td>185.0</td>
<td>181.7</td>
<td>-</td>
<td>-3.5</td>
</tr>
<tr>
<td>Et-Nap</td>
<td>193.9</td>
<td>190.0</td>
<td>182.6</td>
<td>-</td>
<td>-3.9</td>
</tr>
<tr>
<td>i-Pr-Nap</td>
<td>198.0</td>
<td>193.7</td>
<td>184.1</td>
<td>-</td>
<td>-4.3</td>
</tr>
<tr>
<td>HO-Nap</td>
<td>178.1</td>
<td>-</td>
<td>176.4</td>
<td>142.2 (-OH)</td>
<td>-1.7</td>
</tr>
<tr>
<td>MeO-Nap</td>
<td>191.0</td>
<td>187.3</td>
<td>185.2</td>
<td>164.6 (-OMe)</td>
<td>-3.7</td>
</tr>
<tr>
<td>EtO-Nap</td>
<td>195.3</td>
<td>192.7</td>
<td>188.0</td>
<td>171.4 (-OEt)</td>
<td>-2.6</td>
</tr>
<tr>
<td>CN-Nap</td>
<td>194.8</td>
<td>135.5</td>
<td>141.3</td>
<td>198.1 (-NC)</td>
<td>3.3</td>
</tr>
<tr>
<td>NH$_2$-Nap</td>
<td>211.1</td>
<td>200.1</td>
<td>193.7</td>
<td>204.4 (-NH$_2$)</td>
<td>-7.0</td>
</tr>
</tbody>
</table>

$^a$ Experimental Ag$^+$ affinities, $\Delta H_0$, determined by the standard kinetic method as listed in Table 5.1.

$^b$ Theoretical Ag$^+$ affinities, $\Delta H_0$, calculated at the CCSD(T)/[HW(f), 6-31+G(d)] level using the MP2/[HW, 3-21G(d)] optimized geometries. The affinities for the most stable binding mode are shown in bolded font.

$^c$ Theoretical affinities for Ag$^+$ binding to the benzene-$\pi$ ring of substituted naphthalene with substituent attached.

$^d$ Theoretical affinities for Ag$^+$ binding to the benzene-$\pi$ ring of substituted naphthalene without substituent attached.

$^e$ Theoretical affinities for Ag$^+$ binding to the O/N heteroatom binding site of the substituent (in parenthesis).

$^f$ Deviation of theoretical affinities (for the most stable binding mode) from the experimental affinities.
5.2.3 Ag (I) Cation-π Interaction: Ag (I) Cation Bound Complexes of Alkyl and Alkoxy Substituted Naphthalenes

As discussed earlier, cation-π is the energetically preferred Ag⁺ binding mode for alkyl and alkoxy substituted naphthalenes. Similar to alkylbenzenes (Chapter 4), the Ag⁺ cation-π binding to X-Nap is mainly electrostatic in nature, though charge-transfer covalent interaction is also present. Natural population analysis (NPA) with MP2/3-21G(d) wavefunctions reveal that the Ag⁺ ion tends to retain most of its positive charge (≈ 0.91 − 0.92) in these complexes. Furthermore, the Ag⁺-cation-π binding distances (≈ 2.2 − 2.3 Å) of the Ag⁺-(X-Nap) complexes are much longer than the covalent bonds (C-C) length (≈ 1.45 − 1.55 Å), indicating that the Ag⁺-(X-Nap) bonding is far from purely covalent.

![Graph](image)

**Figure 5.8** Plot of Ag⁺ affinities of substituted naphthalenes (X-Nap) versus substituted benzenes (X-Bz) (where X = alkyl substituents (○) and alkoxy substituents (●)).
A good correlation line ($R^2 \sim 0.99$) in Figure 5.8 reveals that the Ag$^+$ binding affinities of the substituted naphthalenes and substituted benzenes (Table A5.1, Appendix II) are induced by similar substituent effects. The Ag$^+$ affinities of both X-Bz and X-Nap were found to increase with the chain length of both alkyl and alcoxy substituents. Given the very similar nature of the benzene and naphthalene rings, one expects that the underlying factors affecting the binding affinities described in the benzene chapter (Section 4.2.3) would be applicable to naphthalene and substituted naphthalene systems. Here, we will only highlight some key trends between the two aromatic systems so that the influence of the fused benzene rings (versus a single benzene ring) on cation-$\pi$ interaction can be elucidated.

5.2.4 Influence of Extended Network on Cation-$\pi$ Interaction

The effect of the additional benzene ring of naphthalene on cation-$\pi$ interaction can be examined by comparing the Ag$^+$ binding affinities of substituted naphthalenes to that of analogous substituted benzenes as shown in Figure 5.8 and Table A5.1. As can be seen from Table 5.4, X-Nap ($X = \text{Me, Et, i-Pr, -OH, -OME, -OEt}$) binds more strongly than X-Bz in all cases. The increase in Ag$^+$ cation-$\pi$ binding energies of X-Nap, relative to those of X-Bz, varies between 15.4 and 20.2 kJ mol$^{-1}$. This corresponds to an average increase in the strength of Ag$^+$ binding of $\sim 11\%$.

**Ion-Quadrupole Interaction.** The theoretical quadrupole moment ($Q_{zz}$) of substituted naphthalenes, calculated by Miss H. M. Lee at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level, are summarized in Table 5.4. The quadrupole moments calculated for substituted naphthalenes are found to be more negative than substituted benzenes with
the same substituent attached. The trend of decreasing molecular quadrupole moments with increasing chain length in alkyl substituted benzene (Section 4.2.3) is also observed in the alkyl substituted naphthalenes, thus reinforcing that the interaction between Ag\(^+\) and the alkyl substituted benzene and naphthalene series are indeed closely related.

<table>
<thead>
<tr>
<th>X-Nap</th>
<th>(Q_{zz} \times 10^{40} \text{ C m}^2)</th>
<th>(\alpha \times 10^{-24} \text{ cm}^3)</th>
<th>(\mu)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nap</td>
<td>-38.5 (-45)</td>
<td>15.0 (17.5)</td>
<td>0.00 (0)</td>
</tr>
<tr>
<td>Me-Nap</td>
<td>-38.1</td>
<td>17.1 (19.52)</td>
<td>0.41</td>
</tr>
<tr>
<td>Et-Nap</td>
<td>-37.5</td>
<td>18.9 (21.36)</td>
<td>0.38</td>
</tr>
<tr>
<td>i-Pr-Nap</td>
<td>-37.5</td>
<td>20.7</td>
<td>0.38</td>
</tr>
<tr>
<td>HO-Nap</td>
<td>-38.1</td>
<td>15.9</td>
<td>1.28</td>
</tr>
<tr>
<td>MeO-Nap</td>
<td>-41.8</td>
<td>17.9</td>
<td>1.42</td>
</tr>
<tr>
<td>EtO-Nap</td>
<td>-41.5</td>
<td>19.8</td>
<td>1.38</td>
</tr>
<tr>
<td>CN-Nap</td>
<td>-1.7</td>
<td>18.1</td>
<td>4.43</td>
</tr>
<tr>
<td>NH(_2)-Nap</td>
<td>-55.1</td>
<td>16.8 (19.73)</td>
<td>2.21</td>
</tr>
</tbody>
</table>

\(a\) Theoretical quadrupole moment, \(Q_{zz}\), calculated at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level of theory, in units of \(10^{40} \text{ C m}^2\). Experimental quadrupole moments were determined by induced-birefringence method (measured in tetrachloromethane (CCl\(_4\)) at 298 K) [Vrbancici and Ritchie, 1980].

\(b\) Theoretical molecular polarizability, \(\alpha\), calculated at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level of theory, in units of \(10^{-24} \text{ cm}^3\). Experimental molecular polarizabilities taken from [Lide, 2002-2003] are shown in the parenthesis.

\(c\) Theoretical molecular dipole moment, \(\mu\), calculated at the MP2/3-21G(d) level of theory, in units of debye D. Experimental molecular dipole moment taken from [Lide, 2002-2003] is shown in the parenthesis.

In the case of alkoxy substitution, a slightly larger quadrupole moment than alkyl substitution is found. Hydroxy and alkoxy substituents are known to have \(\pi\)-electron
donating (via positive resonance or mesomeric effect) properties to the aromatic ring, which are due to ‘delocalization’ of the lone pair electrons on the oxygen atom to the aromatic ring, with corresponding increase in the \( \pi \)-electron density (Scheme 5.1). A stronger resonance effect is found with increasing chain length of the alkyl group of the alkoxy groups. This is probably due to the electron donating properties of the alkyl group, which help to stabilize the partial positive charge on the oxygen atom in the resonance structures. The lone pair electrons on the oxygen atom then become more available to be donated to the aromatic-\( \pi \) ring.

On average, from substituted benzene to substituted naphthalene, the molecular quadrupole moment increases by about 45\% – 72\%. Given the variation in affinity is only about 11\%, it indicates that even though ion-quadrupole interaction has increased from substituted benzene to substituted naphthalene, other unfavorable or destabilizing factors are also present.

![Scheme 5.1 Positive resonance effect of alkoxy substituent (–OR, where R = H, Me, Et).](image)

**Ion-Induced Dipole Interaction.** Experimental molecular polarizabilities (\( \alpha \)) of substituted naphthalenes are not commonly available in the literature. The theoretical molecular polarizabilities, calculated by Miss H. M. Lee at the B3LYP/6-31G(d,p)//MP2/3-21G(d) level, can reproduced the experimental trends for both
substituted benzenes and naphthalenes as shown in Table 4.3 and 5.4. As shown in Figure 5.9, an increasing trend of experimental Ag$^+$ affinities with the theoretical molecular polarizabilities of alkyl and alkoxy substituted naphthalenes is observed, suggesting that the enhanced Ag$^+$ aromatic-π binding to substituted naphthalene should result mainly from increasing ion-induced dipole interaction.

![Graph showing Ag$^+$ affinities vs. theoretical molecular polarizabilities](image)

**Figure 5.9** Plot of Ag$^+$ affinities versus theoretical calculated molecular polarizabilities ($\alpha$, $\times 10^{-24}$ cm$^3$) of X-Nap, where X = alkyl substituents (○) and alkoxy substituents (●).

Again, the increase in Ag$^+$ affinity (~11%) of substituted naphthalenes is relatively small when compared to the significant increases in molecular polarizabilities (~56% – 82%). Given the similar magnitude of change in both the ion-quadrupole and ion-induced dipole interactions, the much smaller change in Ag$^+$ binding affinity of substituted naphthalenes remains unaccounted for.
**Ion-Dipole Interaction.** Like benzene, naphthalene possesses a centre of symmetry and therefore has no permanent molecular dipole moment. Substitution of one of the hydrogen atoms of naphthalene introduces a dipole moment resulting from breaking up of the symmetry in the molecule, regardless of the nature of the substituent.

![Diagram of molecular dipole moments](image)

Figure 5.10 Direction of molecular dipole moments in (a) 2-methoxy-naphthalene (MeO-Nap), (b) 2-cyanonaphthalene (CN-PhOH) and (c) 2-aminonaphthalene (NH₂-Nap).

As shown in Table 5.4, the molecular dipole moments of alkyl substituted naphthalenes are very small in the range of 0 to 0.41 Debye (D). By replacing alkyl with alkoxy substituents, the molecular dipole moments (ranging from 1.28 to 1.42 D) increase about 2 – 3 times because of a larger difference of electronegativity across C and O (2.20, 2.55, 3.44 for H, C and O, respectively [Allred, 1961]). However, as discussed in Section 4.2.3, the direction of the dipole moment is also important for effective ion-dipole interaction apart from its magnitude. As a typical example, the direction of the molecular dipole moment vector of MeO-Nap is shown in Figure 5.10a. Similarly, the molecular dipole moments of alkyl and alkoxy substituted naphthalenes lie in the plane of the aromatic ring, while the Ag⁺ ion interacts with the aromatic-π ring in a direction perpendicular to the plane. Therefore, ion-dipole
interactions should have only minor effects on the strength of Ag\(^+\) cation-π binding for alkyl and alkoxy substituted naphthalenes.

*Nuclear-Nuclear / Electron-Electron Repulsion.* The significant changes in ion-quadrupole and ion-induced dipole interaction is not reflected quantitatively in the Ag\(^+\) binding affinity. This suggests that from benzene to naphthalene, some unfavorable repulsive interaction has increased along the favorable ion-quadrupole and ion-induced dipole attractive interactions. We attribute this to nuclear-nuclear and electron-electron repulsion. Naphthalene, when compared to benzene, has four more carbon and two hydrogen atoms (i.e. 26 nuclear protons and 26 electrons). The additional nuclear protons / electrons would interact repulsively with the nucleus / electrons on the Ag\(^+\) ion, leading to a smaller than expected increase in Ag\(^+\) binding affinity.

5.2.5 Competition between Cation-π and Non-π Interactions: Ag (I) Cation Bound Complexes of 2-Cyanonaphthalene and 2-Aminonaphthalene

Unlike alkyl and alkoxy substituted naphthalenes, Ag\(^+\) prefers to bind to non-π heteroatom nitrogen (N) binding sites in 2-cyanonaphthalene (CN-Nap) and 2-aminonaphthalene (NH\(_2\)-Nap). Similar to cation-π complexes, the binding of Ag\(^+\) ion to these nitrogen sites is also mainly electrostatic in nature; NPA analysis indicate that 0.86 – 0.93 of the unit charges remain with the Ag\(^+\) ion. The underlying causes for the preferred Ag\(^+\) binding to the nitrogen site differ from one substituent to another, as could be revealed by the correlation of Ag\(^+\) affinities with molecular quadrupole moment, dipole moment and polarizability of substituted naphthalenes shown in Table 5.4.
Scheme 5.2 Negative resonance effect of (a) nitro (−NO$_2$), (b) cyano (−CN) substituent, and (c) positive resonance effect of amino (−NH$_2$) substituent.

In the case of the cyano (−CN) substituent, it exerts a negative resonance effect (Scheme 5.2b), which withdraws π-electrons away from the aromatic ring and lowers the molecular quadrupole moment (−1.7 × 10$^{-40}$ C m$^2$). On the other hand, a comparatively strong molecular dipole moment is found in CN-Nap (4.43 D) due to the presence of the electronegative nitrogen atom in the substituent. Consequently, Ag$^+$ prefers to bind to the −CN functional group, in near-perfect alignment with the molecular dipole moment, and in the same plane of the aromatic ring (Figure 5.7 and 5.10b). With reduced π electron density / quadrupole moment but increase in dipole moment, Ag$^+$ binding to the aromatic-π ring is destabilized in CN-Nap, and Ag$^+$
binding to the nitrogen of the cyano substituent is 56.8 – 62.6 kJ mol⁻¹ more stable than the Ag⁺-aromatic π binding for CN-Nap (Table 5.3). Thus, the enhanced Ag⁺ binding strength of CN-Nap (194.8 kJ mol⁻¹) compared to Nap (176.8 kJ mol⁻¹) is primarily due to the strong ion-dipole and ion-induced dipole interactions present.

Ag⁺ binding to NH₂-Nap is interesting in that this non-π binding mode is preferred over the cation-π binding mode by 4.7 – 10.7 kJ mol⁻¹, even though the amino (–NH₂) substituent is π-electron donating (positive resonance effect, Scheme 5.2c), resulting a very large quadrupole moment (–55.1 × 10⁻⁴⁰ C m²) of NH₂-Nap. The Ag⁺ affinity of NH₂-Nap is the highest among the nine substituted naphthalenes investigated in the present study, although the molecular polarizability of NH₂-Nap (16.83 × 10⁻²⁴ cm³) is only slightly larger than Nap, and the dipole moment of NH₂-Nap (2.21 D, Table 5.4 and Figure 5.10c) is small, and opposite in direction when compared to that of CN-Nap. Unlike the case of CN-Nap, Ag⁺ ion aligns itself with the nitrogen lone pair of the –NH₂ substituent in the Ag⁺-(NH₂-Nap) complex, in a direction nearly perpendicular to the plane of the aromatic ring (Figure 5.7). This preference of Ag⁺ binding to amino nitrogen (e.g. NH₃ and alkylamines) over O-donor sites (e.g. H₂O, CH₃OH, (CH₃)₂O, (CH₃)₂CO, etc.) has been noted in previous theoretical and experimental studies [Ma, 1998; El Aribi et al., 2002a; 2002b]. For example, the Ag⁺ affinity [Ma, 1998] of NH₃ is significantly greater (~40 to 50 kJ mol⁻¹) than that of H₂O, despite the fact that water shows a greater dipole moment (1.85 D) than NH₃ (1.47 D). In the case of NH₂-Nap, Ag⁺····NH₂– binding is stronger than Ag⁺ cation-π binding, even though NH₂-Nap has a very large quadrupole moment. This suggests that apart from ion-quadrupole interaction, local Ag⁺ interaction with the nitrogen lone-pair of electrons is important in enhancing.
significantly the ion-induced dipole and charge-transfer interactions in the Ag\(^+\) binding to the N-amino site of NH\(_2\)-Nap. The enhanced charge-transfer interaction is attributed to the small energy gap between the donor and acceptor orbitals, as reflected in the comparable ionization energies of Ag (7.58 eV, [Lide, 2002-2003]) and NH\(_2\)-Nap (7.55 eV, [Lias, et al., 2003]) [Dargel et al., 1999].

Comparing the -CN / -NH\(_2\) substituted naphthalene with that of benzene, the molecular polarizabilities have increased quite modestly for CN-Nap (~ 40%) and NH\(_2\)-Nap (~ 42%). Thus, the increase in Ag\(^+\) affinity for CN-Nap and NH\(_2\)-Nap in the non-\(\pi\) Ag\(^+\) binding mode would be attributed, at least partially, to the ion-induced dipole interaction. What is interesting is the competition between \(\pi\) and non-\(\pi\) mode of Ag\(^+\) binding as a function of ring size. The difference in \(\pi\) and non-\(\pi\) mode affinity in -CN and -NH\(_2\) substituted benzene is 79 and 17 kJ mol\(^{-1}\), respectively, in favor of the non-\(\pi\) mode. This preference decreased to 57 and 11 kJ mol\(^{-1}\), respectively, in the substituted napthalene analogues. In other words, the \(\pi\) mode has become more competitive against the non-\(\pi\) mode in the fused benzene system. We note that the calculated quadrupole moment of CN-Nap (-1.7 × 10\(^{-40}\) C m\(^2\)) is less negative than that of benzonitrile (CN-Bz, -2.4 × 10\(^{-40}\) C m\(^2\)), suggesting that ion-quadrupole interaction has become smaller. However, given the small numerical value of the quadrupole moment and the associated theoretical error bars involved, the change is small, so that, to first approximation, the ion-quadrupole contribution to the \(\pi\) mode of binding is fairly constant from CN-Bz to CN-Nap. Since the increase in dipole moment is associated with favoring the non-\(\pi\) mode of binding, one may conclude that the increase in molecular polarizability is the key
factor behind the increase in competitiveness of the \( \pi \) mode of binding for fused benzene ring systems.

5.2.6 Comparison of Ag (I) Cation Affinities with Other Metal Cations

Affinities of Naphthalene

Studies on cation-\( \pi \) interactions of different metal cations with naphthalene (Nap) have been very limited. The experimental and theoretical binding affinities of naphthalene available in the literature are summarized in Table 5.5.

Table 5.5 Experimental binding affinities of different metal cations (\( M^+ \)) to naphthalene at 0 K (\( \Delta H_0 \) in kJ mol\(^{-1} \))

<table>
<thead>
<tr>
<th>( M^+ )</th>
<th>( \Delta H_0 ) (Expt)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag(^+)</td>
<td>176.8(^a)</td>
</tr>
<tr>
<td>Li(^+)</td>
<td>187.2 ± 16.4(^b)</td>
</tr>
<tr>
<td>Na(^+)</td>
<td>107.1 ± 5.0(^b)</td>
</tr>
<tr>
<td>K(^+)</td>
<td>80.9 ± 5.1(^b)</td>
</tr>
<tr>
<td>Rb(^+)</td>
<td>73.0 ± 4.9(^b)</td>
</tr>
<tr>
<td>Cs(^+)</td>
<td>69.3 ± 5.4(^b)</td>
</tr>
<tr>
<td>Fe(^+)</td>
<td>&gt; 174(^c)</td>
</tr>
</tbody>
</table>

\(^a\) Our work as shown in Table 5.3.
\(^b\) Threshold-CID, \( \Delta H_0 \) [Arunugama and Rodgers, 2003a].
\(^c\) Radiative association kinetic analysis, [Dunbar et al., 1994].

Similar to benzene (Bz) and toluene (Tol), Arunugama and Rodgers found that the binding strength of Nap with alkali metal decrease from Li\(^+\) to Cs\(^+\) with an increasing ionic size, and the trend has been interpreted on purely electrostatic
grounds [Amunugama and Rodgers, 2003a]. Likewise, our experimental Ag⁺ affinity of Nap (176.8 kJ mol⁻¹) is greater than that of Na⁺, and is exceptional to this trend. As in the case of alkylbenzenes, this can be accounted for as due to the presence of enhanced covalent interaction in the case of Ag⁺ (Section 4.2.5).

The affinity of Fe⁺-Nap has been determined by Dunbar and his co-workers using the radiative association reaction method [Dunbar et al., 1994], with a lower limit of 174 kJ mol⁻¹. Theoretical studies have shown that Fe⁺-Bz binding is primarily electrostatic in nature, but with a substantive contribution from charge-transfer due to back donation of metal 3d electrons to the π* orbitals of Bz [Bauschicher et al., 1992]. Furthermore, the large binding strength in Fe⁺-Nap is a result from sd-hybridization, which can reduce the Pauli repulsion and further stabilizes the Fe⁺-Nap complex [Rodger et al., 2000; Amunugama and Rodgers, 2003a]. Given that the Fe⁺-Bz affinity is lower than that of Ag⁺-Bz affinity (previous chapter), the lower bound value (> 174 kJ mol⁻¹) obtained by Dunbar is at least qualitatively consistent with the Ag⁺-Nap affinity determined here (176.8 kJ mol⁻¹).
Chapter 6  Ag (I) Cation Affinities of Substituted Indoles

6.1  Background

Indole is made up of a fused benzene and pyrrole ring (refer to Figure 1.2c), and can be considered as the model cation-π ligand for the amino acid tryptophan. There are two possible cation-π binding sites in indole: Ag⁺ binding to the benzene-π or pyrrole-π aromatic ring. Non-π Ag⁺ binding to the pyrrole nitrogen of indole is also possible.

Theoretical studies on Na⁺, K⁺, Mg⁺, Al⁺, and Cr⁺ binding to indole are reported recently [Mecozz et al., 1996; Dunbar, 1998; Ryzhov and Dunbar, 1999]. However, there is only one experimental study on metal cation binding to indole in the gas phase. The experimental Na⁺, K⁺, Mg⁺, Al⁺, Ca⁺, Cr⁺, Fe⁺ and Mo⁺ affinities of indole were determined radiative association kinetics in a Fourier transform-ion cyclotron resonance mass spectrometer [Dunbar et al., 1994].

The present study is undertaken to measure the Ag⁺ affinities of substituted indoles (X-indole, where X = H, –Me, –OH, –OMe, –NO₂ and –CN), and to evaluate the effect of these substituents on the relative stabilities of the cation-π and non-π binding modes.
6.2 Results and Discussion

6.2.1 Determination of Ag (I) Cation Affinities of Substituted Indoles

Relative Ag (I) Cation Affinities of Alkyl and Alkoxy Substituted Indoles.

Following the same procedures mentioned in previous chapters, the relative Ag$^+$ affinities of 10 substituted indoles including indole, 1-methylindole (1-Me-Indole), 2-methylindole (2-Me-Indole), 3-methylindole (3-Me-Indole), 5-methylindole (5-Me-Indole), 6-methylindole (6-Me-Indole), 7-methylindole (7-Me-Indole), 5-hydroxyindole (5-HO-Indole), 5-methoxyindole (5-MeO-Indole), 7-methoxyindole (7-MeO-Indole), were investigated under low-energy Ar-CID conditions at two different collision energies of 20 eV and 40 eV, as well as under high-energy (4.7 keV, laboratory frame) He-CID conditions. Typical examples of low-energy Ar-CID (20 eV and 40 eV, laboratory frame) and high-energy (4.7 keV, laboratory frame) He-CID mass spectra of a Ag$^+$ bound heterodimer, [Indole + Ag + (6-Me-Indole)]$^+$ (where 6-Me-Indole = 6-methylindole), are shown in Figure 6.1.

Similar to the case of alkylbenzenes and substituted naphthalenes, the relative Ag$^+$ affinities of isobaric substituted indoles (same relative molar mass), such as 1-Me-Indole and 2-Me-Indole, were determined from their respective heterodimers with other substituted indoles. The relative Ag$^+$ affinity ladders are internally consistent, and summarized in Figure 6.2 and A6.1 (Appendix III).

As shown in Figure 6.2 and A6.1, the experimental uncertainty (SD) of the \( \ln\left([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+] \right) \) ratios are \( \sim 0.06 \) (n = 5, or \( \pm 0.15 \) at 90 % confidence interval) between each successive stair step in the relative Ag$^+$ affinity ladders. This means that the low-energy CID affinity ladder measurements cannot distinguish small
differences in Ag⁺ affinities in the order of ± 2 kJ mol⁻¹. Under such circumstances, the order of Ag⁺ affinities of the ten substituted indoles were found to be:


A similar order of Ag⁺ affinities was found from the high-energy CID Ag⁺ affinity ladder shown Figure A6.2 (Appendix III).

**Absolute Ag (I) Cation Affinities of Alkyl and Alkoxy Substituted Indoles.** In order to obtain the absolute Ag⁺ affinities, the conversion factor 1/RTₑff in Equation 2.10 has to be found by the same approach as described in Chapter 5, i.e. with reference to a set of reference alkylbenzenes of known Ag⁺ affinities under identical experimental conditions. Typical examples of low-energy Ar-CID mass spectra of the Ag⁺ bound heterodimer, [Indole + Ag + (m-Xy)]⁺, are shown in Figure A6.3 (Appendix III). Their ln([Ag + L₁]⁺ / [Ag + L₂]⁺) ratios are also included in Figure 6.2 and A.6.1, and are found to be internally consistent.

The calibration plots of cumulative ln([Ag + L]⁺ / [Ag + Bz]⁺) values obtained under low-energy Ar-CID conditions at collision energies of 20eV (Figure 6.2) and 40eV (Figure A6.1, Appendix III) versus the Ag⁺ affinities of RBz at 0K, including Bz, Tol, m-Xy, 1,3,5-Me₃Bz, are shown in Figure 6.3 and A6.4 (Appendix III), respectively. The Tₑff values derived from their slope are 1,206 K and 1,637 K, respectively, and the absolute Ag⁺ affinities can then be found by extrapolating of the linear plots with respect to the reference alkylbenzenes affinity values. The results, together with the experimental uncertainty of the linear regression plots (Figure 6.3 and A6.4) are summarized in Table 6.1.
Figure 6.1  Low-energy ((a) 20eV and (b) 40eV, laboratory frame) Ar-CID and (c) high-energy (4.7 keV, laboratory frame) He-CID mass spectra of Ag (I) cation bound heterodimer, [Indole + Ag + (6-Me-Indole)]\(^+\) where 6-Me-Indole = 6-methylindole).
<table>
<thead>
<tr>
<th>Ligands, L</th>
<th>$\ln([Ag + L_1]^<em>/[Ag + L_2]^</em>)$</th>
<th>$\ln([Ag + L]^<em>/[Ag + Bz]^</em>)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bz</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$m$-Xy</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1,3,5-Me$_3$Bz</td>
<td>2.98</td>
<td>1.73</td>
</tr>
<tr>
<td>5-HO-Indole</td>
<td>2.51</td>
<td>1.49</td>
</tr>
<tr>
<td>Indole</td>
<td>2.06</td>
<td>1.29</td>
</tr>
<tr>
<td>5-MeO-Indole</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3-Me-Indole</td>
<td>1.32</td>
<td>0.62</td>
</tr>
<tr>
<td>5-Me-Indole</td>
<td>1.03</td>
<td>0.33</td>
</tr>
<tr>
<td>6-Me-Indole</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7-Me-Indole</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1-Me-Indole</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2-Me-Indole</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7-MeO-Indole</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Figure 6.2** Experimentally measured $\ln([Ag + L_1]^*/[Ag + L_2]^*)$ values for the low-energy (20 eV, laboratory frame) Ar-CID of Ag$^+$ bound heterodimers of substituted indoles ($T_{\text{eff}} = 1,206$ K). The $\ln([Ag + L_1]^*/[Ag + L_2]^*)$ values are the logarithm of ion intensity ratios. The data presented under the heading The $\ln([Ag + L]^*/[Ag + Bz]^*)$ are average ± standard deviation of cumulative values expressed relative to benzene (Bz). The values in ( ) are the arithmetic difference of measured values having a common reference ligands.
The Ag⁺ affinities of alkyl and alkoxy substituted indoles obtained at the two collision energies (20 eV and 40 eV) are in very good agreement with each other; the mean absolute deviation is only 2.2 kJ mol⁻¹. Such good agreement lends further confidence to the quantitative results obtained by the standard kinetic method in this study. The average of these two sets of Ag⁺ affinities was adopted as our final result. Combining the experimental uncertainty (SD, ~3 kJ mol⁻¹, Figure 6.3) with the estimated error of the reference Ag⁺ affinity values (10 – 11 kJ mol⁻¹, Section 4.2.2), the overall uncertainty of the experimentally measured Ag⁺ affinities of alkyl and alkoxy substituted indoles is ± 13 – 14 kJ mol⁻¹.

Figure 6.3  Calibration plot of cumulative ln([Ag + L]⁺ / [Ag + Bz]⁺) values (Figure 6.2) obtained under low-energy (20eV, laboratory frame) Ar-CID conditions (T_{eff} = 1,206 K) versus Ag⁺ affinities of reference alkylbenzenes of Bz, Tol, m-Xy, 1,3,5-Me3Bz (○) at 0K. The absolute Ag⁺ affinities of alkyl and alkoxy substituted indoles (○) were determined from interpolation and extrapolation of the plots.
<table>
<thead>
<tr>
<th>X-Indole</th>
<th>Low-energy (20 eV, lab frame) Ar-CID&lt;sup&gt;a&lt;/sup&gt;</th>
<th>Low-energy (40 eV, lab frame) Ar-CID&lt;sup&gt;b&lt;/sup&gt;</th>
<th>Average&lt;sup&gt;c&lt;/sup&gt;</th>
</tr>
</thead>
<tbody>
<tr>
<td>Indole</td>
<td>207.1 ± 1.7 (3.7)</td>
<td>207.5 ± 0.1 (0.2)</td>
<td>207.3 ± 1.7</td>
</tr>
<tr>
<td>1-Me-Indole</td>
<td>220.6 ± 2.3 (5.0)</td>
<td>223.2 ± 0.1 (0.2)</td>
<td>221.9 ± 2.3</td>
</tr>
<tr>
<td>2-Me-Indole</td>
<td>222.2 ± 2.4 (5.1)</td>
<td>224.6 ± 0.1 (0.2)</td>
<td>223.4 ± 2.4</td>
</tr>
<tr>
<td>3-Me-Indole</td>
<td>213.4 ± 2.0 (4.3)</td>
<td>215.0 ± 0.1 (0.2)</td>
<td>214.2 ± 2.0</td>
</tr>
<tr>
<td>5-Me-Indole</td>
<td>214.4 ± 2.1 (4.4)</td>
<td>215.9 ± 0.1 (0.2)</td>
<td>215.1 ± 2.1</td>
</tr>
<tr>
<td>6-Me-Indole</td>
<td>215.5 ± 2.1 (4.5)</td>
<td>217.9 ± 0.1 (0.2)</td>
<td>216.7 ± 2.1</td>
</tr>
<tr>
<td>7-Me-Indole</td>
<td>218.0 ± 2.2 (4.7)</td>
<td>220.9 ± 0.1 (0.2)</td>
<td>219.4 ± 2.2</td>
</tr>
<tr>
<td>5-HO-Indole</td>
<td>199.4 ± 1.4 (2.9)</td>
<td>202.0 ± 0.1 (0.1)</td>
<td>200.7 ± 1.4</td>
</tr>
<tr>
<td>5-MeO-Indole</td>
<td>212.0 ± 1.9 (4.1)</td>
<td>214.4 ± 0.1 (0.2)</td>
<td>213.2 ± 1.9</td>
</tr>
<tr>
<td>7-MeO-Indole</td>
<td>223.1 ± 2.5 (5.2)</td>
<td>226.7 ± 0.1 (0.2)</td>
<td>224.9 ± 2.5</td>
</tr>
</tbody>
</table>

<sup>a</sup> Experimental Ag<sup>+</sup> affinities, ΔH<sub>0</sub>, measured by low-energy (20eV, laboratory frame) Ar-CID calibration plot (Figure 6.3, T<sub><sup>e</sup></sub><sup>et</sup>= 1,206 K). The uncertainties are given as ± SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot.

<sup>b</sup> Experimental Ag<sup>+</sup> affinities, ΔH<sub>0</sub>, measured by low-energy (40eV, laboratory frame) Ar-CID calibration plot (Figure A6.4, Appendix III, T<sub><sup>e</sup></sub><sup>et</sup>= 1,637 K). The uncertainties given as ± SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot.

<sup>c</sup> Average ± SD of two low-energy Ar-CID values, with standard deviation obtained from the equation s<sup>2</sup> = s<sub>1</sub><sup>2</sup> + s<sub>2</sub><sup>2</sup>, where s<sub>1</sub><sup>2</sup> and s<sub>2</sub><sup>2</sup> are the SD of low-energy Ar-CID 20 eV (laboratory frame) and 40 eV (laboratory frame) values.

**Absolute Ag (I) Cation Affinities of 5-Nitroindole and 5-Cyanoindole.** Apart from the alkyl and alkoxy substituted indoles (cation-π binding ligands) investigated in the previous sections, the Ag<sup>+</sup> binding affinity of 5-nitroindole (5-NO<sub>2</sub>-Indole) and 5-cyanoindole (5-CN-Indole) were also measured. Since entropic effects are present in the dissociation of the Ag<sup>+</sup> bound heterodimers between 5-NO<sub>2</sub>-Indole / 5-CN-Indole and alkyl / alkoxy-substituted indoles, the Ag<sup>+</sup> affinity of 5-CN-Indole and 5-NO<sub>2</sub>-
Indole have to be measured by the extended kinetic method using alkylbenzenes (1,3,5-Me₂Bz, 1,2,4,5-Me₄Bz, Me₃Bz and Me₄Bz) as reference compounds.

Examples of the low-energy Ar-CID mass spectra of the heterodimer, [(5-NO₂-Indole) + Ag + (m-Xy)]⁺ and [(5-CN-Indole) + Ag + Me₅Bz]⁺, subjected to various collision energies (in terms of eV), are shown in Figure 6.4 and A6.5 (Appendix III), respectively. The corresponding first plot and second plots of extended kinetic method measurements of (5-NO₂-Indole) + Ag + RBz]⁺ and [(5-CN-Indole) + Ag + RBz]⁺ heterodimers are shown in Figure 6.5 and A6.6 (Appendix III), respectively. The Ag⁺ affinity of 5-NO₂-Indole and 5-CN-Indole, as well as, the Δ(ΔS_Ag⁺)_{app} term, are summarized in Table 6.2. Combining the maximum uncertainty (SD) from the extended kinetic plots (0.3 kJ mol⁻¹, Table 6.2) with the estimated error of the reference Ag⁺ affinities (10 – 11 kJ mol⁻¹, Section 4.2.2), the experimental Ag⁺ affinities of 5-NO₂-Indole and 5-CN-Indole have an uncertainty (SD) of ±11 – 12 kJ mol⁻¹.

<table>
<thead>
<tr>
<th>X-Indole</th>
<th>ΔH₀ (Exp) (^a)</th>
<th>Δ(ΔS_Ag⁺)_{app}</th>
</tr>
</thead>
<tbody>
<tr>
<td>5-NO₂-Indole</td>
<td>196.6 ± 0.3 (0.6)</td>
<td>4.0 ± 0.2 (0.3)</td>
</tr>
<tr>
<td>5-CN-Indole</td>
<td>213.5 ± 0.2 (0.3)</td>
<td>-0.7 ± 0.1 (0.3)</td>
</tr>
</tbody>
</table>

\(^a\) Experimental affinities, ΔH₀, determined by the extended kinetic method as shown in Figure 6.5 and A6.6 (Appendix III). Weighted results with experimental uncertainties given as ± SD (90% confidence interval) in the format of [Armentrout, 2000].
**Figure 6.4** Triple quadrupole MS/MS spectrum of the [(5-NO₂-Indole) + Ag + (m-Xy)]⁺ heterodimer using argon as collision gas at a various collision energies (laboratory frame) (5-NO₂-Indole = 5-nitroindole and m-Xy = m-xylene).
Figure 6.5  (a) Plot of \( \ln(\text{[Ag + (5-NO₂-Indole)]}^+ / \text{[Ag + RBz]}^+) \) versus \( |\Delta H_{\text{Ag} + \text{RBz}}^+ - \Delta H_{\text{Avg}}| \) at different collision energies and (b) plot of \( [\Delta G_{\text{Ag} + \text{(5-NO₂-Indole)}}^\text{app} - \Delta H_{\text{Avg}}] / \text{RT}_{\text{eff}} \) versus \( 1/\text{RT}_{\text{eff}} \) for the heterodimers \( (5-\text{NO₂-Indole}) + \text{Ag + RBz}^+ \) (5-NO₂-Indole = 5-nitroindole and RBz = Tol, m-Xy, 1,2,4,5-Me₄Bz, Me₃Bz). The Ag⁺ affinity of 5-NO₂-Indole, \( \Delta H_{\text{Ag} + \text{(5-NO₂-Indole)}}^+ \), and the entropy changes, \( \Delta(\Delta S_{\text{Ag}}^+) \text{app} \), are expressed as unweighted values.
6.2.2 Comparison of Experimental and Theoretical Ag (I) Cation Affinities of Substituted Indoles

The Ag⁺ binding affinities at 0 K calculated at the CCSD(T)/[HW(f), 6-31+G(d)] level and based on optimized geometries obtained at the MP2/[HW, 3-21G(d)] level are summarized in Table 6.3. Their most stable structures are shown in Figure 6.6. Other stable binding modes / geometries with lower Ag⁺ affinities for 5-HO-Indole, 7-MeO-Indole and 5-NO₂-Indole are shown in Figure A6.7 to A6.9 (Appendix III).

Similar to substituted benzenes and naphthalenes, Ag⁺ binds strongly to the π faces for alkyl and alkoxy substituted indoles (either the benzene-π or pyrrole-π binding, as shown in Figure 6.6). The benzene-π and pyrrole-π binding modes are very similar in their Ag⁺ binding affinities (small differences ~ 1 – 11 kJ mol⁻¹), and are generally more stable (on average of ~20 kJ mol⁻¹) than the non-π binding, except for 5-NO₂-Indole and 5-CN-Indole, for which Ag⁺ binding to the O/N heteroatom(s) in the non-π binding mode is preferred (Figure 6.6). For 5-NO₂-Indole, the Ag⁺ ion binds bidentately to the two oxygen atoms of the nitro substituent. In case of 5-CN-Indole, the Ag⁺ ion binds to the nitrogen of the –CN substituent. For 5-NO₂-Indole and 5-CN-Indole, the non-π binding energies are about 40 kJ mol⁻¹ higher (more stable) than their respective cation-π binding energies.

As shown in Table 6.3, our experimental values are in fairly good agreement with theoretical values for the most stable binding Ag⁺ modes, with an average deviation of ± 9 kJ mol⁻¹, although the theoretical affinity values are systematic lower than the experimental values except for 5-NO₂-Indole. Considering the relatively large size of Ag⁺-substituted indole systems, such theoretical values could be regarded as very acceptable.
Qualitatively, the same order of theoretical and experimental affinity values within Me-Indole, O-Indole series, N-Indole series are found:

(I) Me-Indole Series: \(1\text{-Me-Indole} \sim 2\text{-Me-Indole}) > (3\text{-Me-Indole} \sim 5\text{-Me-Indole} \sim 6\text{-Me-Indole} \sim 7\text{-Me-Indole}) > (\text{Indole})

(II) O-Indole Series: \((7\text{-MeO-Indole}) > (5\text{-MeO-Indole}) > (5\text{-HO-Indole})

(III) N-Indole Series: \((5\text{-CN-Indole}) > (5\text{-NO}_2\text{-Indole})

Upon methyl and methoxy substitution, the \(\text{Ag}^+\) cation-\(\pi\) binding affinities are increased relative to indole. With the exception of 3-Me-Indole, methyl substitution at the pyrrole-\(\pi\) ring (i.e. 1 and 2-positioned Me-Indole) leads to a larger \(\text{Ag}^+\) binding strength than methyl substitution at the benzene-\(\pi\) ring (i.e., 5, 6 and 7-positioned Me-Indole). On the other hand, the difference in affinity values, both experimental and theoretical, are small \((-2\text{ kJ mol}^{-1}\)) for the 1-Me/2-Me, 7-Me/6-Me and 6-Me/5-Me pairs such that their theoretical and experimental order of affinities are in disagreement. This could be due to the combined errors of the experimental measurements and the theoretical calculations, as is also found in the case of Indole / 5-OH-Indole pair. Given the small difference in \(\text{Ag}^+\) binding energies of the pyrrole-\(\pi\) and benzene-\(\pi\) binding modes, another plausible cause is that \(\text{Ag}^+\) binding to pyrrole-\(\pi\) and benzene-\(\pi\) are occurring under low-energy CID conditions, leading to experimental affinities not strictly following the theoretical order of the most stable pyrrole-\(\pi\) or benzene-\(\pi\) binding modes.
Figure 6.6 Optimized geometries of the most stable $\text{Ag}^+$-(X-Indole) complexes at MP2/[HW, 3-21G(d)] level of theory. The distance (in Å) from $\text{Ag}^+$ to the centre of the aromatic ring that lies in the plane of the carbon atoms, and $\theta$ represents the angel of deviation (in degrees) of $\text{Ag}^+$ to the vertical line perpendicular to the center of the ring. Only selected bond lengths and angles are presented.
Table 6.3  Experimental and theoretical Ag⁺ affinities, ΔH₀, at 0K (in kJ mol⁻¹) of substituted indoles

<table>
<thead>
<tr>
<th>X-Indole</th>
<th>ΔH₀ (Exp)</th>
<th>Benzene-π</th>
<th>Pyrrole-π</th>
<th>Pyrrole-N</th>
<th>Heteratom at Substituent</th>
<th>Δ(ΔH₀ (Theory)) - Δ(ΔH₀ (Exp))</th>
<th>ΔH₀ (Theory) - ΔH₀ (Exp)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Indole</td>
<td>207.3</td>
<td>195.2</td>
<td>198.8</td>
<td>164.4</td>
<td>-</td>
<td>0.0</td>
<td>-8.5</td>
</tr>
<tr>
<td>1-Me-Indole</td>
<td>221.9</td>
<td>203.1</td>
<td>213.4</td>
<td>-</td>
<td>-</td>
<td>0.1</td>
<td>-8.5</td>
</tr>
<tr>
<td>2-Me-Indole</td>
<td>223.4</td>
<td>205.2</td>
<td>212.8</td>
<td>-</td>
<td>-</td>
<td>-2.0</td>
<td>-10.6</td>
</tr>
<tr>
<td>3-Me-Indole</td>
<td>214.2</td>
<td>201.4</td>
<td>-</td>
<td>176.7</td>
<td>-</td>
<td>-4.3</td>
<td>-12.8</td>
</tr>
<tr>
<td>5-Me-Indole</td>
<td>215.1</td>
<td>206.1</td>
<td>205.1</td>
<td>-</td>
<td>-</td>
<td>-0.5</td>
<td>-9.0</td>
</tr>
<tr>
<td>6-Me-Indole</td>
<td>216.7</td>
<td>204.9</td>
<td>205.5</td>
<td>-</td>
<td>-</td>
<td>-2.6</td>
<td>-11.2</td>
</tr>
<tr>
<td>7-Me-Indole</td>
<td>219.4</td>
<td>204.5</td>
<td>203.4</td>
<td>-</td>
<td>-</td>
<td>-6.4</td>
<td>-14.9</td>
</tr>
<tr>
<td>5-HO-Indole</td>
<td>200.7</td>
<td>194.6</td>
<td>199.1</td>
<td>166.1</td>
<td>174.5</td>
<td>7.0</td>
<td>-1.6</td>
</tr>
<tr>
<td>5-MeO-Indole</td>
<td>213.2</td>
<td>206.7</td>
<td>206.2</td>
<td>-</td>
<td>188.6 (-OMe)</td>
<td>2.1</td>
<td>-6.5</td>
</tr>
<tr>
<td>7-MeO-Indole</td>
<td>224.9</td>
<td>210.2</td>
<td>210.6</td>
<td>204.4 (Pyrrole-N &amp; -OMe)</td>
<td>-5.8</td>
<td>-14.3</td>
<td></td>
</tr>
<tr>
<td>5-NO₂-Indole</td>
<td>196.6</td>
<td>171.9</td>
<td>160.5</td>
<td>125.0</td>
<td>206.3 (-O₂N)</td>
<td>-</td>
<td>9.7</td>
</tr>
<tr>
<td>5-CN-Indole</td>
<td>213.5</td>
<td>154.0</td>
<td>165.1</td>
<td>130.5</td>
<td>212.8 (-NC)</td>
<td>-</td>
<td>-0.7</td>
</tr>
</tbody>
</table>

a Experimental Ag⁺ affinities, ΔH₀, determined by standard kinetic method as shown in Table 6.1.
b Experimental Ag⁺ affinities, ΔH₀, determined by extended kinetic method as shown in Table 6.2.
c Theoretical Ag⁺ affinities, ΔH₀, at the CCSD(T)/[HW(f), 6-31+G(d)] level by using the MP2/[HW, 3-21G(d)] optimized geometries. The affinities with the most stable geometry are shown in bolded font.
d Theoretical affinities for Ag⁺ binding to the benzene-π ring of substituted indole as shown in Figure A6.7 (Appendix III).
e Theoretical affinities for Ag⁺ binding to the pyrrole-π ring of substituted indole as shown in Figure A6.7.
f Theoretical affinities for Ag⁺ binding to the nitrogen heteroatom at the pyrrole ring of substituted indole as shown in Figure A6.7.
g Theoretical affinities for Ag⁺ binding to the oxygen or nitrogen heteroatom(s) of the substituent (in parenthesis) as shown in Figure A6.7.
h Difference between theoretical (most stable binding mode) and experimental relative affinities, expressed with reference to Indole.
i Difference between theoretical (most stable binding mode) and experimental absolute affinities.
6.2.3 Ag (I) Cation-π Interaction: Ag (I) Cation Bound Complexes of Alkyl and Alkoxy Substituted Indoles

Similar to X-Nap (Chapter 5), Ag⁺ cation-π binding to alkyl and alkoxy substituted indoles is mainly electrostatic in nature. Natural population analysis (NPA) shows that the Ag⁺ ion tends to retain most of its positive charge (≈ 0.84 – 0.92) upon complexation to alky and alkoxy substituted indoles. Furthermore, the Ag⁺-XIndole binding distances (≈ 2.2 – 2.7 Å) were also much longer than C-C covalent bonds length (≈ 1.45 – 1.55 Å), indicating that Ag⁺-(X-Indole) binding is definitely not purely covalent. Given that naphthalene is a 10 π-electron homocyclic analogue of indole, it is expected that the basic factors affecting the binding affinities described previously in naphthalene and substituted naphthalenes (Chapter 5) would be applicable to indole and substituted indoles. Here we will discuss some special features found in Ag⁺-substituted indole complexes.

Ag⁺-(5-NO₂-Indole) Benzene-π Binding Complex. An interesting feature is noted in the Ag⁺-(NO₂-Indole) complex in the cation-π (benzene-π) binding mode; the Ag⁺ affinities is unexpectedly larger. Indeed in this complex, the Ag⁺ is not oriented centrally but tilted to interact with the oxygens of the nitro-substituent (Figure A6.9, Appendix III). There are several possible reasons to account for the extraordinary large affinity. Firstly, the large molecular dipole moment (μ = 6.47 D, Figure 6.7c) boosts up the binding strength through the ion-dipole interaction. Additionally, it is likely that the Ag⁺ ion is involved in binding to a five-member ring (consisting of the partial double bond character of N-O and C-C) structure and the Ag⁺-complex is somehow stabilized.
Figure 6.7  Direction of molecular dipole moments in (a) 1-methylindole (1-Me-Indole), (b) 3-methylindole (3-Me-Indole), (c) 5-methylindole (5-Me-Indole), (d) 7-methoxyindole (7-MeO-Indole), (e) 5-nitroindole (5-NO₂-Indole) and (f) 5-cyanoindole (5-CN-Indole).

Effect of Methyl Substitution Position on Benzene-\(\pi\) and Pyrrole-\(\pi\) Binding Mode.

As commented earlier, the energy difference in the two \(\pi\) binding modes are small. However, it appears that the methyl substituent tends to have a larger effect when it is substituted into the pyrrole ring, but not much effect when it is substituted into benzene ring. For example, when the methyl substituent is in 1- and 2-positions (i.e. on the pyrrole ring), the Ag\(^{+}\) affinities in the pyrrole-\(\pi\) binding mode are \(\approx\)10 kJ mol\(^{-1}\) larger than the benzene-\(\pi\) binding mode, but when the methyl group is substituted into benzene ring at the 5-, 6- and 7-positions, the Ag\(^{+}\) affinities differ by no more than 1 kJ mol\(^{-1}\). This suggests that, due to the resonance delocalization of the nitrogen lone pair of electrons, the pyrrole ring has greater \(\pi\)-electron density than
the benzene ring in substituted indoles, leading to stronger Ag$^+$ binding to the pyrrole ring.

6.2.4 Cation-π Interaction in Indole: the Influence of the Pyrrole Nitrogen Heteroatom

Indole is a composite of a benzene-π ring and a pyrrole-π ring. Different from other simple π faces like benzene and naphthalene, the nitrogen present in the pyrrole ring has a larger electronegativity than carbon, leading to a disturbance in the symmetry of the π-electron system. The effect of the additional pyrrole-π ring, as well as the electronegativity of the nitrogen present in this extended π system, can be elucidated by comparing the Ag$^+$ binding affinities of X-Indole to those of X-Bz and X-Nap obtained previously.

The Ag$^+$ affinities of substituted indoles are higher than their substituted benzene and naphthalene analogs. Compared to substituted benzenes, the presence of an additional pyrrole ring increases the experimental Ag$^+$ affinities by 26 – 33% (19 – 27% in theoretical affinities). This could be attributed to the larger molecular quadrupole moment and polarizabilities of the substituted indoles.

Ion-Quadrupole Interaction. The quadrupole moment moments of alkyl and alkoxy substituted indoles (Table 6.4), ranging from -39.3 to -50.7 × 10^{-40} C m^2, are much larger more (negative) than those X-Bz and X-Nap with the same substituent attached, as exemplified by the order: indole (-42.3 × 10^{-40} C m^2) > naphthalene (-38.5 × 10^{-40} C m^2) > benzene (-24.2 × 10^{-40} C m^2). On average, from substituted
benzene to substituted indole, the molecular quadrupole moment increases by about 70 – 92 %, but the maximum increase in Ag⁺ affinity is only about 33 %. This suggests that nuclear-nuclear and electron-electron repulsion are more significant in substituted indoles.

<table>
<thead>
<tr>
<th>X-Indole</th>
<th>Q_{zz}</th>
<th>α</th>
<th>μ</th>
</tr>
</thead>
<tbody>
<tr>
<td>Indole</td>
<td>-42.3</td>
<td>12.5</td>
<td>2.26</td>
</tr>
<tr>
<td>1-Me-Indole</td>
<td>-39.3</td>
<td>14.4</td>
<td>2.63</td>
</tr>
<tr>
<td>2-Me-Indole</td>
<td>-40.1</td>
<td>14.7</td>
<td>2.58</td>
</tr>
<tr>
<td>3-Me-Indole</td>
<td>-40.9</td>
<td>14.3</td>
<td>2.14</td>
</tr>
<tr>
<td>5-Me-Indole</td>
<td>-44.1</td>
<td>14.5</td>
<td>2.02</td>
</tr>
<tr>
<td>6-Me-Indole</td>
<td>-42.3</td>
<td>14.5</td>
<td>2.12</td>
</tr>
<tr>
<td>7-Me-Indole</td>
<td>-39.4</td>
<td>14.3</td>
<td>2.43</td>
</tr>
<tr>
<td>5-HO-Indole</td>
<td>-44.6</td>
<td>13.3</td>
<td>2.37</td>
</tr>
<tr>
<td>5-MeO-Indole</td>
<td>-50.7</td>
<td>15.2</td>
<td>1.86</td>
</tr>
<tr>
<td>7-MeO-Indole</td>
<td>-47.5</td>
<td>15.0</td>
<td>2.27</td>
</tr>
<tr>
<td>5-NO₂-Indole</td>
<td>-17.2</td>
<td>15.9</td>
<td>6.47</td>
</tr>
<tr>
<td>5-CN-Indole</td>
<td>-11.5</td>
<td>15.4</td>
<td>6.68</td>
</tr>
</tbody>
</table>


a Theoretical Quadrupole moment, Q_{zz}, calculated at the B3LYP/6-31G(d,p)/MP2/3-21G(d) level of theory, in units of 10⁻⁴⁰ C m².

b Theoretical molecular polarizability, α, calculated at the B3LYP/6-31G(d,p)/MP2/3-21G(d) level of theory, in units of 10⁻²⁴ cm³.

c Theoretical molecular dipole moments, μ, calculated at the MP2/3-21G(d) level of theory, in units of debye D.
**Ion-Induced Dipole Interaction.** As shown in Table 6.4, the molecular polarizabilities of methyl, hydroxyl and methoxy substituted indoles vary over a small range of $13.3 - 15.2 \times 10^{-24}$ cm$^3$, which could explain, at least partly, the narrow range of Ag$^+$ affinities (200.7 – 224.9 kJ mol$^{-1}$) of these substituted indoles. The theoretical molecular polarizabilities of substituted indoles are 21 – 52% greater than their substituted benzene analogs (Table 6.4 and Table A5.1), which is not matched by the comparatively smaller increase (26 – 33%) in experimental Ag$^+$ affinities. This confirms again that nuclear-nuclear or electron-electron repulsion is indeed greater in aromatic indole-$\pi$ systems.

**Ion-Dipole Interaction.** Unlike Bz and Nap, indole possesses a non-zero permanent molecular dipole moment (calculated to be 2.26 Debye, Table 6.4), which is directed away from the nitrogen across the center of the molecule to the opposite end of the benzene ring, corresponding to the flow of $\pi$-electrons away from the nitrogen lone pair (Figure 6.7). As commented earlier, the ion-dipole interaction is not a determining factor in cation-$\pi$ interaction of X-Bz and X-Nap complexes due to the small dipole moment and ineffective orientation lying in a plane perpendicular to the Ag$^+$ ion. However, the situation in the case of Ag$^+$-(X-Indole) complexes is completely different. A more accurate measure of the contribution of ion-dipole interaction is the effective dipole moment ($\mu_{\text{eff}}$, refer to Footnote 6.1). The $\mu_{\text{eff}}$ of indole, HO-, Me-, and MeO-substituted indoles in benzene-$\pi$ binding are 0.15 – 0.30 D Å$^{-2}$ (as compared to 0.04 – 0.13 D Å$^{-2}$ for X-Bz, 0.02 – 0.12 D Å$^{-2}$ for X-Nap, 0.35 D Å$^{-2}$ for H$_2$O), while in pyrrole-$\pi$ binding are only 0.07 – 0.16 D Å$^{-2}$. The molecular dipole moment vectors are not exactly lying in a plane perpendicular to the Ag$^+$ ion. Therefore, apart from molecule quadrupole moment and polarizability,
ion-dipole interaction may also contribute to the larger Ag⁺ cation-π affinities found in substituted indoles, and the effect of dipole moment in benzene-π binding is larger than the pyrrole-π binding. The dipole moment effect is also noted in previous study of metal cation binding to indole [Dunbar, 1998; Gapeev et al., 2000].

Footnote 6.1 Effective dipole moment ($\mu_{\text{eff}}$ in the unit of Debye Å²) is a more accurate description of the strength of ion-dipole interaction by taking the angle of alignment into consideration. The effective dipole moment is defined as:

$$\mu_{\text{eff}} = \mu \cdot \cos(\Phi) / r_{\mu}^2$$

where $\mu = \text{molecular dipole moment of the deformed ligand in the complexed state}$

(from Mulliken population analysis, in Debye)

$\Phi = \text{angles of deviation between the cation and the dipole moment vector (in degrees)}$

$r_{\mu} = \text{the distance between the cation and the center of the dipole moment vector (in Å)}$
Comparison between Substituted Indoles and Substituted Naphthalenes. The experimental and theoretical Ag⁺ cation-π affinities of X-Indole are found to be 12 – 19% and 8 – 15% greater than the corresponding X-Nap cation-π complexes. These percentages are approximately equal to half of the energy difference between the Ag⁺-(X-Indole) and Ag⁺-(X-Bz) complexes. Although the Ag⁺ affinities of these substituted indole complexes are larger, their molecular polarizabilities are 14 – 17% smaller than substituted naphthalene. Given that the molecular polarizabilities are smaller, the higher Ag⁺ affinity of X-Indole should be contributed by other factors such as enhanced quadrupole moment and dipole moment. In actual practice, the quadrupole moments of these substituted indoles are found to be 3 – 17% greater than the corresponding substituted naphthalenes. Apart from the quadrupole moment, the dipole moment factor should also play a vital role in enhancing the Ag⁺ affinities of substituted indoles.

Based on the Ag⁺ affinities order: X-Indole > X-Nap > X-Bz, we could conclude that addition of a 5-member heterocyclic (nitrogen-containing) ring is more effective in enhancing Ag⁺ cation-π binding than that of 6-member homocyclic ring.

6.2.5 Competition between Cation-π and Non-π Interactions: Ag (I) Cation Bound Complexes of 5-Nitroindole and 5-Cyanoindole

As in the case in X-Nap, Ag⁺ prefers to bind to non-π O/N heteroatom binding sites in 5-nitroindole (5-NO₂-Indole) and 5-cyanoindole (5-CN-Indole). The factors affecting the choice of non-π binding site in 5-CN-Indole is the same as in CN-Nap discussed previously (Section 5.2.5).
Like cyano (−CN) substituent, the nitro (−NO$_2$) substituent is also well known for its π-electron withdrawing property (negative resonance effect, Scheme 5.2a), leading to reduced π electron density in the aromatic ring / quadrupole moment (−17.2 × 10$^{-40}$ C m$^2$), and destabilization of Ag$^+$ cation-π binding. On the other hand, the Ag$^+$ binding to the non-π binding sites of 5-NO$_2$-Indole and 5-CN-Indole is much stronger due to more effective ion-dipole interaction (Figure 6.7f and 6.7e).

5-NO$_2$-Indole has slightly greater dipole moment and molecular polarizability than 5-CN-Indole, but its experimental and theoretical Ag$^+$ affinities at 196.6 and 206.3 kJ mol$^{-1}$ are lower than the 5-CN-Indole by 16.9 and 6.5 kJ mol$^{-1}$, respectively. This is also indicated by the longer Ag$^+$⋯O bonding distance (≈2.38 Å) in Ag$^+$-(5-NO$_2$-Indole) and the shorter Ag$^+$⋯N bonding distance (≈2.17 Å) in Ag$^+$-(5-CN-Indole). The higher Ag$^+$ affinity of 5-CN-Indole could be attributed to the preference of Ag$^+$ binding to N-donor over O-donor sites [Ma, 1998].

6.2.6 Comparison of Ag (I) Cation Affinities with Other Metal Cations

Affinities of Indole

Very few experimental metal cation affinities of indole have been reported. The available experimental and theoretical metal cation binding affinities of indole are summarized in Table 6.5. As discussed in Section 6.2.2, the most stable binding site of the Ag$^+$-Indole complex is at the pyrrole-π ring. Unlike Ag$^+$ Ryzhov and Dunbar found that Na$^+$, K$^+$, Mg$^+$, Al$^+$, Cr$^+$ prefer to bind to the benzene-π site [Dunbar, 1998; Ryzhov and Dunbar, 1999].
Table 6.5  Experimental and theoretical metal cations (M⁺) binding affinities of indole at 0 K (ΔH₀ in kJ mol⁻¹)

<table>
<thead>
<tr>
<th>M⁺</th>
<th>ΔH₀ (Exp)</th>
<th>ΔH₀ (Theory)</th>
<th>Best-estimate ΔH°</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag⁺</td>
<td>207.3a</td>
<td>198.8a</td>
<td>–</td>
</tr>
<tr>
<td>Na⁺</td>
<td>137.7b</td>
<td>127.2c</td>
<td>142.3 ± 12.6</td>
</tr>
<tr>
<td></td>
<td></td>
<td>146.4d</td>
<td></td>
</tr>
<tr>
<td>K⁺</td>
<td>104.6b</td>
<td>87.4c</td>
<td>104.6 ± 12.6</td>
</tr>
<tr>
<td>Mg⁺</td>
<td>200.8b</td>
<td>171.1c</td>
<td>192.5 ± 16.7</td>
</tr>
<tr>
<td></td>
<td></td>
<td>213.4d</td>
<td></td>
</tr>
<tr>
<td>Al⁺</td>
<td>200.8b</td>
<td>168.2c</td>
<td>188.3 ± 16.7</td>
</tr>
<tr>
<td></td>
<td></td>
<td>192.5d</td>
<td></td>
</tr>
<tr>
<td>Cr⁺</td>
<td>&gt; 200.8b</td>
<td>188.7c</td>
<td>196.6 ± 16.7</td>
</tr>
<tr>
<td>Fe⁺</td>
<td>&gt; 200.8b</td>
<td>–</td>
<td>&gt; 200.8</td>
</tr>
</tbody>
</table>

a  Our work as shown in Table 6.3.
b  Radiative association kinetic analysis, [Ryzhov and Dunbar, 1999].
c  DFT calculation at B3LYP level of theory, ΔH₀, [Ryzhov and Dunbar, 1999].
d  DFT calculation at B3LYP level of theory, ΔH₀, [Dunbar, 1998].
e  Best estimated absolute binding energy based on available experimental data supplemented by quantum chemistry comparative estimates, [Ryzhov and Dunbar, 1999].

Both our experimental and theoretical Ag⁺ absolute affinities of indole are larger than the Na⁺ and K⁺ values determined by the radiative association kinetic analysis, B3LYP calculations, as well as the ‘best-estimated’ values. As discussed previously, this is due to the greater charge-transfer (covalent) interaction found in Ag⁺-cation-π interactions.

The reported experimental Mg⁺, Al⁺, Cr⁺ and Fe⁺ affinities for indole (~ 200 kJ mol⁻¹) are very close to our Ag⁺ affinity at 207.3 kJ mol⁻¹. Given the relatively large experimental uncertainties (12-17 kJ mol⁻¹) reported for these affinity measurements, it is premature to conclude on the order or trend of indole binding affinities among transitional metals.
Chapter 7  Ag (I) Cation Affinities of Substituted Phenols

7.1  Background

Phenol is the model ligand representing the side chain of the naturally occurring aromatic amino acid tyrosine (Tyr). It has two potential cation binding sites: the aromatic π-ring and the hydroxy oxygen. The binding interactions of Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Mg⁺, Al⁺, Cr⁺, Fe⁺ and Co⁺ to phenol have been previously reported [Mecoazzi et al., 1996; Ryzhov and Dunbar, 1999; Armentrout and Rodgers; 2000; Hoyau et al., 1999; McMahon and Ohanessian; 2000; Amunugama and Rodgers, 2002c; Schroeter et al., 1998], but the Ag⁺ binding affinities of phenol and substituted phenols have not been investigated.

For substituted phenols (X-PhOH), the situation is even more complex as the two substituent groups (–OH and –X) on the benzene ring can exert their influences on different possible binding modes and Ag⁺ binding affinities. For example, the –OH substituent can act as a π-electron donor via the resonance effect, but the –OH binding site could also be affected by the electron-donating/withdrawing property of the substituent X and vice versa. In this study, the effect of different substituents of X-PhOH, where X = H, Me, Et, i-Pr, t-Bu, OMe, OEt, NO₂ at the para-positioned, on Ag⁺ binding affinities and the preference for cation-π or non-π modes of binding in substituted phenol are investigated.
7.2 Results and Discussion

7.2.1 Determination of Ag (I) Cation Affinities of Substituted Phenols

Relative Ag (I) Cation Affinities of Alkyl and Alkoxy Substituted Phenols. Similar to the case of substituted indoles, the [(X-PhOH$_1$) + Ag + (X-PhOH$_2$)]$^+$ heterodimers generated from seven para-substituted phenols, including phenol (PhOH), 4-methylphenol (Me-PhOH), 4-ethylphenol (Et-PhOH), 4-$i$-propylphenol ($i$-Pr-PhOH), 4-$t$-butylphenol ($t$-Bu-PhOH), 4-methoxyphenol (MeO-PhOH) and 4-ethoxyphenol (EtO-PhOH), were monitored by low-energy Ar-CID at two different levels of collisional activation (different collision energies at 25eV and 40eV). Typical examples of low-energy (25 eV and 40 eV, laboratory frame) Ar-CID mass spectra of Ag$^+$ bound heterodimer, [PhOH + Ag + (Et-PhOH)]$^+$, are shown in Figure 7.1. The relative Ag$^+$ affinity ladders are found to be internally consistent (Figure 7.2 and Figure A7.1 in Appendix IV). The order of Ag$^+$ affinities among the seven substituted phenols are:

PhOH < Me-PhOH ~ MeO-PhOH < Et-PhOH ~ EtO-PhOH < $i$-Pr-PhOH < $t$-Bu-PhOH.

Absolute Ag (I) Cation Affinities of Alkyl and Alkoxy Substituted Phenols. Under the same circumstances as described previously, the absolute Ag$^+$ affinities of alkyl and alkoxy substituted phenols were established by ‘inclusion’ of reference alkylbenzenes including benzene (Bz), toluene (Tol) and $m$-xylene ($m$-Xy) into the relative affinity ladders (Figure 7.2 and A7.1). Typical examples of low-energy Ar-CID mass spectra of the Ag$^+$ bound heterodimer, [(MeO-PhOH) + Ag + Bz]$^+$, are shown in Figure A7.2 (Appendix IV).
Figure 7.1  Low-energy (a) 25eV and (b) 40eV, laboratory frame) Ar-CID mass spectra of Ag (I) cation bound heterodimer, [PhOH + Ag + (Et-PhOH)]\(^+\), (PhOH = phenol and Et-PhOH = 4-ethylphenol).

Calibration plots of the cumulative \( \ln([\text{Ag} + L]^+ / [\text{Ag} + \text{Bz}]^+) \) values obtained under low-energy Ar-CID condition at collision energies of 25eV (Figure 7.2) versus the Ag\(^+\) affinities of reference RBz are shown in Figure 7.3. The \( T_{\text{eff}} \) values obtained from the slopes of the plots are 1,092 K (for alky substituted phenols, Figure 7.3a) and 1,100 K (for alkoxy substituted phenols, Figure 7.3b), respectively. The absolute Ag\(^+\) affinities at 0 K of alky and alkoxy substituted phenols were determined from interpolation and extrapolation of the plots (Equation 2.10). Similarly, the calibration plots obtained at 40 eV (laboratory frame) were
constructed (Figure A7.3, Appendix IV), with the $T_{\text{eff}}$ values determined to be 1,448 K and 1,552 K, respectively. The absolute Ag$^+$ affinities of alkyl and alkoxy substituted phenols obtained in these effective temperatures are summarized in Table 7.1.

![Graph with data](image)

**Figure 7.2** Experimentally measured $ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ values for the low-energy (25 eV, laboratory frame) Ar-CID of Ag$^+$ bound heterodimers of (a) alkyl substituted phenols ($T_{\text{eff}} = 1,092$ K) (b) alkoxy substituted phenols ($T_{\text{eff}} = 1,100$ K). The $ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ values are the logarithm of ion intensity ratios. The data presented under the heading $ln([\text{Ag} + L]^+ / [\text{Ag} + \text{Bz}]^+)$ are average ± standard deviation of cumulative values expressed relative to benzene (Bz).
Figure 7.3  Calibration plots of cumulative $\ln([\text{Ag} + \text{L}]^+ / [\text{Ag} + \text{Bz}]^+)$ values (Figure 7.2) obtained under low-energy (25eV, laboratory frame) Ar-CID ($T_{\text{eff}}$ = (a) 1,092 K and (b) 1,100 K) conditions versus Ag$^+$ affinities of reference alkylbenzenes of Bz, Tol, m-Xy (●) at 0K. The absolute Ag$^+$ affinities of alkyl and alkoxy substituted phenols (○) were determined from interpolation and extrapolation of the plots.
Table 7.1  Experimental $\text{Ag}^+$ affinities, $\Delta H_0$, at 0K (in kJ mol$^{-1}$) of substituted phenols

<table>
<thead>
<tr>
<th>X-PhOH</th>
<th>$\Delta H_0$ (Expt)</th>
<th>Low-energy (25 eV lab frame) Ar-CID$^a$</th>
<th>Low-energy (40 eV lab frame) Ar-CID$^b$</th>
<th>Average$^c$</th>
</tr>
</thead>
<tbody>
<tr>
<td>PhOH</td>
<td>157.9 ± 0.1 (0.2)</td>
<td>158.7 ± 0.5 (1.0)</td>
<td>158.3 ± 0.5</td>
<td></td>
</tr>
<tr>
<td>Me-PhOH</td>
<td>165.0 ± 0.1 (0.2)</td>
<td>165.3 ± 0.4 (0.8)</td>
<td>165.2 ± 0.4</td>
<td></td>
</tr>
<tr>
<td>Et-PhOH</td>
<td>171.7 ± 0.1 (0.1)</td>
<td>172.8 ± 0.2 (0.8)</td>
<td>172.2 ± 0.4</td>
<td></td>
</tr>
<tr>
<td>i-Pr-PhOH</td>
<td>178.3 ± 0.1 (0.2)</td>
<td>179.3 ± 0.5 (1.1)</td>
<td>178.8 ± 0.5</td>
<td></td>
</tr>
<tr>
<td>t-Bu-PhOH</td>
<td>182.9 ± 0.1 (0.3)</td>
<td>184.4 ± 0.7 (1.4)</td>
<td>183.7 ± 0.7</td>
<td></td>
</tr>
<tr>
<td>MeO-PhOH</td>
<td>165.9 ± 0.1 (0.1)</td>
<td>165.4 ± 0.4 (0.8)</td>
<td>165.6 ± 0.4</td>
<td></td>
</tr>
<tr>
<td>EtO-PhOH</td>
<td>171.0 ± 0.1 (0.1)</td>
<td>170.9 ± 0.4 (0.8)</td>
<td>170.9 ± 0.4</td>
<td></td>
</tr>
</tbody>
</table>

$^a$ Experimental $\text{Ag}^+$ affinities, $\Delta H_0$, determined from low-energy (25 eV, laboratory frame) Ar-CID calibration plots ($T_{\text{eff}} = 1,092 \text{ K}$ for alkyl substituted phenols, Figure 7.3a or $T_{\text{eff}} = 1,100 \text{ K}$ for alkoxy substituted phenols, Figure 7.3b). The uncertainties are given as $\pm$ SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plots (Figure 7.3).

$^b$ Experimental $\text{Ag}^+$ affinities, $\Delta H_0$, determined from low-energy (40 eV, laboratory frame) Ar-CID calibration plots ($T_{\text{eff}} = 1,448 \text{ K}$ for alkyl substituted phenols, Figure A7.3a or $T_{\text{eff}} = 1,552 \text{ K}$ for alkoxy substituted phenols, Figure A7.3b). The uncertainties are given as $\pm$ SD (90% confidence interval) of the linear regression analysis [Meier and Zund, 2000] of the calibration plot (Figure A7.3, Appendix IV).

$^c$ Average $\pm$ SD of two low-energy Ar-CID values, with standard deviation obtained from the equation $s^2 = s_1^2 + s_2^2$, where $s_1$ and $s_2$ are the SD of low-energy Ar-CID 25 eV (laboratory frame) and 40 eV (laboratory frame) values.

As shown in Table 7.1, excellent agreement was found between the $\text{Ag}^+$ affinities of substituted phenols obtained under two different collision energies (25 eV and 40 eV, laboratory frame), with a mean absolute deviation (MAD) of 0.8 kJ mol$^{-1}$. The average of these two sets of $\text{Ag}^+$ affinities is adopted as our final results. Combining the error of the affinity values of reference ligands ($10 - 11$ kJ mol$^{-1}$, Section 4.2.2) and the maximum uncertainty (SD) of the linear regression plots (0.7 kJ mol$^{-1}$, Figure 7.3 and A7.3), the overall experimental uncertainty (SD) is $\pm 11 - 12$ kJ mol$^{-1}$. 
**Absolute Ag (I) Cation Affinities of 4-Nitrophenol.** As discussed later (Section 7.2.2 and refer to Figure 7.6), the Ag⁺ ion binds to the two oxygen atoms of the -NO₂ substituent in 4-nitrophenol, and is different from the cation-π binding modes of alkyl and alkoxy substituted phenols. Hence, NO₂-PhOH could not be included into the relative affinity ladder of alkyl and alkoxy substituted phenols (Figure 7.2 and A7.1) due to the presence of entropic effects.

<table>
<thead>
<tr>
<th>X-PhOH</th>
<th>ΔH₀ (Exp)ᵃ</th>
<th>Δ(ΔS₄⁺) app</th>
</tr>
</thead>
<tbody>
<tr>
<td>NO₂-PhOH</td>
<td>166.5 ± 0.7 (1.2)</td>
<td>1.0 ± 0.5 (1.0)</td>
</tr>
</tbody>
</table>

ᵃ Experimental Ag⁺ affinities, ΔH₀, determined by the extended kinetic method as shown in Figure 7.8. Weighted results with experimental uncertainties given as ± SD (90% confidence interval) in the format of [Armentrout, 2000].

Therefore, the Ag⁺ affinity of 4-nitrophenol has to be determined by the extended kinetic method. Typical low-energy Ar-CID mass spectra of the heterodimer, [(NO₂-PhOH) + Ag + Bz]⁺, subjected to various collision energies (eV, laboratory frame), are shown in Figure 7.4. Using the experimental Ag⁺ affinities of alkylbenzenes (RBz, including Bz, Tol, m-Xy and 1,3,5-Me₃Bz) as reference values, the corresponding first and second plots derived from the extended kinetic method measurements are shown in Figure 7.5a and 7.5b, respectively. The measured Ag⁺ affinity of NO₂-PhOH is 166.5 kJ mol⁻¹ (Table 7.2), with the experimental uncertainty of 1.2 kJ mol⁻¹ (90% confidence interval). Combining the error of the affinity values of reference ligands (10 – 11 kJ mol⁻¹, Section 4.2.2), the overall estimated uncertainty is ± 11 – 12 kJ mol⁻¹.
Figure 7.4  Triple quadrupole MS/MS spectrum of the [(NO$_2$-PhOH) + Ag + Bz]$^+$ heterodimer using argon as collision gas at a various collision energies (laboratory frame) (NO$_2$-PhOH = 4-nitrophenol and Bz = benzene).
Figure 7.5  
(a) Plot of \( \ln([\text{Ag} + (\text{NO}_2\text{-PhOH})^+]/[\text{Ag} + \text{RBz}]^+) \) versus \( \Delta H_{[\text{Ag} + \text{RBz}]^+} - \Delta H_{\text{Avg}} \) at different collision energies and (b) plot of \( [\Delta G_{\text{app}}^{\text{app}}] + (\text{NO}_2\text{-PhOH})^+ - \Delta H_{\text{Avg}}] / RT_{\text{eff}} \) versus \( 1/RT_{\text{eff}} \) for the heterodimers \((\text{NO}_2\text{-PhOH}) + \text{Ag} + \text{RBz}^+\) (NO2-PhOH = 4-nitrophenol and RBz = Bz, Tol, m-Xy, 1,3,5-Me3Bz). The Ag\(^+\) affinity of NO2-PhOH, \( \Delta H_{[\text{Ag} + (\text{NO}_2\text{-PhOH})]^+} \), and the apparent entropy change, \( \Delta(\Delta S_{\text{Ag}}^{\text{app}})^+ \), are expressed as unweighted values.
Figure 7.6 Optimized geometries of the most stable Ag⁺-(X-PhOH) complexes at MP2/[HW, 3-21G(d)] level of theory. r represents the distance (in Å) from Ag⁺ to the centre of the aromatic ring that lies in the plane of the carbon atoms, and θ represents the angel of deviation (in degrees) of Ag⁺ to the vertical line perpendicular to the center of the ring. Only selected bond lengths and angles are presented.
7.2.2  Comparison of Experimental and Theoretical Ag (I) Cation Affinities of Substituted Phenols

The theoretical Ag\(^+\) binding energies of PhOH, Me-PhOH, Et-PhOH, MeO-PhOH, EtO-PhOH and NO\(_2\)-PhOH, determined at CCSD(T)/[HW(f), 6-31+G(d)] level of theory, are summarized in Table 7.3, and the most stable Ag\(^+\) binding modes are shown in Figure 7.6. Taking MeO-PhOH and NO\(_2\)-PhOH as examples, the binding geometries of other stable Ag\(^+\)-(X-PhOH) complexes are illustrated in Figure A7.4 and A7.5 (Appendix IV), respectively.

As shown in Table 7.3, Ag\(^+\) preferentially binds to the aromatic-\(\pi\) ring of alkyl or alkoxy substituted phenols, rather than at the hydroxy oxygen (O). The cation-\(\pi\) binding modes of alkyl substituted phenols are 21.6 (for PhOH) to 31.1 (for Et-PhOH) kJ mol\(^{-1}\) more stable than the non-\(\pi\) modes. For alkoxy substituted phenols, Ag\(^+\) binding to the alkoxy oxygen is 8.4 - 16.7 to kJ mol\(^{-1}\) less stable than the cation-\(\pi\) complexes. On the other hand, Ag\(^+\) prefers to bind bidentately to the two oxygen atoms of the –NO\(_2\) substituent of NO\(_2\)-PhOH (non-\(\pi\) mode); the non-\(\pi\) mode is \(-44\) kJ mol\(^{-1}\) and \(-80\) kJ mol\(^{-1}\) more stable than Ag\(^+\) cation-\(\pi\) and monodentate binding to any one of the hydroxy oxygens, respectively.

The experimental and theoretical order of Ag\(^+\) binding affinities is consistent for alkyl and alkoxy substituted phenols, i.e., PhOH < Me-PhOH < Et-PhOH and MeO-PhOH < EtO-PhOH. For PhOH, Me-PhOH and Et-PhOH, the experimental and theoretical affinity values are in very good agreement, with a mean absolute deviation (MAD) of 4.0 kJ mol\(^{-1}\) only (Table 7.3). For MeO-PhOH and EtO-PhOH, the MAD is acceptable but noticeably larger (9.7 kJ mol\(^{-1}\)). Interestingly, a better agreement with MAD at 3.8 kJ mol\(^{-1}\) is obtained if the theoretical values for the
second most stable $\text{Ag}^+$ binding modes (i.e. $\text{Ag}^+$ binds to the alkoxy oxygen) is used in the comparison. We note that the difference in energy between the most stable cation-$\pi$ and second most stable non-$\pi$ binding modes is smaller for the alkoxyphenols (average 12.7 kJ mol$^{-1}$) than the alkylphenols (average 27.3 kJ mol$^{-1}$) (Table 7.3). Thus, it is possible that under our low-energy CID experimental conditions, a mixture of $\text{Ag}^+$ bound $\text{MeO-PhOH}$ and $\text{EtO-PhOH}$ complexes in their most stable cation-$\pi$ and second most stable non-$\pi$ ($\text{Ag}^+$ binding to the alkoxy oxygen) binding modes are present, leading to a lower $\text{Ag}^+$ experimental binding energy measured. In fact, such situation has been suggested to occur in the threshold-CID measurements of alkali metal cation affinities of phenol and methoxybenzene [Arunugama and Rodgers, 2002; 2003]. On the other hand, the MAD between experimental and theoretical values are all within the limits of uncertainty of 11 – 12 kJ mol$^{-1}$ of the kinetic method measurements. Therefore, it is difficult to draw definite conclusions on the occurrence of mixture of $\text{Ag}^+$ binding mode(s) under our experimental conditions.

Furthermore, we noted that a relatively large discrepancy is found between the experimental and theoretical $\text{Ag}^+$ binding affinities of NO$_2$-PhOH; the theoretical value over-estimates the experimental value by 18.0 kJ mol$^{-1}$. This very large discrepancy may arise from the theoretical model employed. In NO$_2$-PhOH, the $\text{Ag}^+$ is bidentately coordinated to the $-\text{NO}_2$ group. However, in the derivation of the protocol, because of the lack of experimental data, only monodentate complexes were included [Ma, 1998]. Clearly, further work is required in order to establish the ability of the theoretical protocol to reproduce the $\text{Ag}^+$ binding affinities of bidentately bound ligands.
At the same time, this discrepancy may be related to the special preference of Ag\(^+\) to form linear di-coordinated ligand complexes [Holland and Castleman, 1982; Deng and Kebarle, 1998; Koizumi et al., 2003]. Holland and Castleman first showed that the enthalpy of binding (in kcal mol\(^{-1}\)), ΔH\(^o\), for Ag(H\(_2\)O)\(_n\)\(^+\) are: ΔH\(^o\)\(_{1,0}\) = 33.3, ΔH\(^o\)\(_{2,1}\) = 24.4, ΔH\(^o\)\(_{3,2}\) = 15.0, ΔH\(^o\)\(_{4,3}\) = 14.9 [Holland and Castleman, 1982], while that of Ag(CH\(_3\)OH)\(_n\)\(^+\) (at 0 K in kJ mol\(^{-1}\)) are: Ag(CH\(_3\)OH)\(^+\) = 152, Ag(CH\(_3\)OH)\(_2\)\(^+\) = 138, Ag(CH\(_3\)OH)\(_3\)\(^+\) = 66, and Ag(CH\(_3\)OH)\(_4\)\(^+\) = 56 [Koizumi et al., 2003]. In other words, the first and second Ag\(^+\) binding enthalpies (affinities) for water and methanol are close to each other, but it drops by 50% or more when Ag\(^+\) is tri- or tetra-coordinated to water or methanol. This phenomenon is in contrast to that observed for alkali ions such as Li\(^+\) and Na\(^+\), where the binding affinity decreases fairly regularly as the coordination number, n, increases [Đžidić and Kebarle, 1970].

The experimental observations were rationalized by Bauschlicher et al. as due to sd\(\sigma\) orbital hybridization of the Ag\(^+\) ion [Bauschlicher et al., 1992].

Briefly, as the first ligand approaches Ag\(^+\), the doubly occupied 4d\(\sigma\) orbital hybridizes with the empty 5s orbital. This sd\(\sigma\) hybridization reduces the electron density along the \(\sigma\) axis, allowing the first and second ligand to approach the Ag\(^+\) more closely in opposite directions (180°) with comparable and mainly electrostatic binding strengths. The large decrease in the third ligand binding energy has been attributed to the loss of sd\(\sigma\) hybridization and to the increasing ligand-ligand repulsion as more ligands are added [Bauschlicher et al., 1991].

We proposed that a similar situation occurs within the Ag\(^+\) binding environment of the dissociating [(NO\(_2\)-PhOH) + Ag + RBz]\(^+\) heterodimer ions used in the kinetic method measurements. If NO\(_2\)-PhOH and alkylbenzenes retain their Ag\(^+\) modes of
binding (Figure 4.3 and 7.6) in the heterodimer ion, the Ag⁺ would be tri-coordinated to two oxygen atoms of NO₂-PhOH and the π-face of alkylbenzenes. Due to the loss of σδ hybridization and electronic repulsion between the electron-rich-NO₂ and Ag⁺, weakening of the binding strength at the Ag⁺......O–N sites and Ag⁺......π surface is expected. This is contrary to the initial assumption of the kinetic method, which presumes that the Ag⁺ binding modes and energies of ligands in the heterodimer is the same as that of the Ag⁺ bound monomer complexes. As a result, the Ag⁺ affinity of nitrophenol measured by the kinetic method is significantly lower than the actual value (Footnote 7.1).

Footnote 7.1 A similar situation may also occur in the case of 5-nitroindole (Table 6.3). While the theoretical Ag⁺ affinity values of other substituted indoles are all lower than the experimental values, the theoretical value of 5-nitroindole is actually 9.7 kJ mol⁻¹ higher than the experimental value, even though the discrepancy between theoretical and experimental values is smaller in the case of 5-nitroindole. The smaller difference could be due to (i) the accuracy of the theoretical protocol varies from one type of ligand to another (e.g. oxygen versus nitrogen-containing ligands), and (ii) the Ag⁺ binding environment of 5-NO₂-Indole in the heterodimer is different from that of the 5-NO₂-PhOH.
Table 7.3  Experimental and theoretical Ag⁺ affinities, \(\Delta H_0\), at 0 K (in kJ mol⁻¹) of substituted phenols

<table>
<thead>
<tr>
<th>X-PhOH</th>
<th>(\Delta H_0) (Exp)ᵃ</th>
<th>Cation-π Bindingᵈ</th>
<th>Hydroxy Oxygen (-OH)ᵉ</th>
<th>Heteroatom of Substituentᶠ</th>
<th>(\Delta H_0) (Theory) - (\Delta H_0) (Exp)ᵍ</th>
</tr>
</thead>
<tbody>
<tr>
<td>PhOH</td>
<td>158.3</td>
<td>159.9</td>
<td>138.3</td>
<td>–</td>
<td>1.6 (-20.0)</td>
</tr>
<tr>
<td>Me-PhOH</td>
<td>165.2</td>
<td>171.7</td>
<td>144.9</td>
<td>–</td>
<td>6.5 (-20.3)</td>
</tr>
<tr>
<td>Et-PhOH</td>
<td>172.2</td>
<td>176.0</td>
<td>144.9</td>
<td>–</td>
<td>3.8 (-27.3)</td>
</tr>
<tr>
<td>i-Pr-PhOH</td>
<td>178.8</td>
<td>–</td>
<td>–</td>
<td>–</td>
<td>–</td>
</tr>
<tr>
<td>t-Bu-PhOH</td>
<td>183.7</td>
<td>–</td>
<td>–</td>
<td>–</td>
<td>–</td>
</tr>
<tr>
<td>MeO-PhOH</td>
<td>165.6</td>
<td>175.7</td>
<td>147.1</td>
<td>159.0 (-OMe)</td>
<td>10.1 (-6.6)</td>
</tr>
<tr>
<td>EtO-PhOH</td>
<td>170.9</td>
<td>180.2</td>
<td>149.5</td>
<td>171.8 (-OEt)</td>
<td>9.3 (0.9)</td>
</tr>
<tr>
<td>NO₂-PhOH</td>
<td>166.5ᵇ</td>
<td>139.9</td>
<td>102.0</td>
<td>184.5 (-O₂N)</td>
<td>18.0 (-26.6)</td>
</tr>
</tbody>
</table>

ᵃ Experimental Ag⁺ affinities, \(\Delta H_0\), determined by standard kinetic method as shown in Table 7.1.
ᵇ Experimental Ag⁺ affinities, \(\Delta H_0\), determined by extended kinetic method as shown in Table 7.2.
ᶜ Theoretical Ag⁺ affinities, \(\Delta H_0\), calculated at the CCSD(T)/[HW(f), 6-31+G(d)] level using MP2/[HW, 3-21G(d)] optimized geometries. The affinities for the most stable binding mode are shown in bold font.
ᵈ Theoretical affinities, \(\Delta H_0\), calculated for Ag⁺ binding to the benzene-π ring of substituted phenols as shown in Figure A7.4 (Appendix IV).
ᵉ Theoretical affinities, \(\Delta H_0\), calculated for Ag⁺ binding to the oxygen atom of the hydroxy group of substituted phenols as shown in Figure A7.4 and A7.5 (Appendix IV).
ᶠ Theoretical affinities, \(\Delta H_0\), calculated for Ag⁺ binding to the oxygen or nitrogen heteroatom binding site of the substituent (in parenthesis) as shown in Figure A7.4 and A7.5 (Appendix IV).
ᵍ Difference between experimental and theoretical affinity for the most stable binding mode. Difference between experimental and theoretical affinity for the second most stable binding modes are shown in parenthesis.
7.2.3 Ag (I) Cation-π Interaction: Ag (I) Cation Bound Complexes of Alkyl and Alkoxy Substituted Phenols

In previous chapters, we have highlighted repeatedly that the interaction between \( \text{Ag}^+ \) and various aromatic-π systems are based mainly on electrostatic interactions. Similar situation was observed in \( \text{Ag}^+-(\text{X-PhOH}) \) complexes, as reflected in the positive charge retained on the \( \text{Ag}^+ \) ion (\( \sim 0.90 - 0.91 \)) calculated by natural population analysis (NPA) at the MP2/3-21G(d) level of theory. Various factors, such as molecular quadrupole moment, dipole moment and polarizability, affect the strength of the electrostatic interaction have been already discussed.

Given the positive resonance effect (i.e., π-electron donor, refer to Scheme 5.1), one may expect the hydroxy substituent would have certain enhancement of cation-π interaction. As shown in previous chapters, hydroxy substitution to Bz and Nap leads to slightly stronger \( \text{Ag}^+ \) cation-π binding strength by 1.2 and 1.3 \( \text{kJ mol}^{-1} \), respectively. The corresponding enhancements in theoretically calculated cation-π affinities are 3.4 and 1.1 \( \text{kJ mol}^{-1} \), respectively. On the contrary, even though the theoretical \( \text{Ag}^+ \) binding energies of alkyl and alkoxy substituted phenols are found to be increased slightly by an average of \( \sim 3.5 \text{kJ mol}^{-1} \) (\( \sim 2\% \)), the experimental binding affinities of alkyl and alkoxy substituted phenols appear to be smaller than those of alkyl and alkoxy substituted benzenes by an average 4.0 \( \text{kJmol}^{-1} \) and 7.5 \( \text{kJ mol}^{-1} \), respectively. This discrepancy may arise from errors of the theoretical protocol and experimental measurements. Such possibilities have already been discussed earlier and will not be repeated here. Hence, we tend to conclude that the X-Bz and X-PhOH employed in our study have comparable \( \text{Ag}^+ \) affinities. The similarity in the \( \text{Ag}^+ \) binding strength of PhOH and Bz should be the result of the
canceling effect of the electronegativity of the hydroxy group against the delocalization of the lone pairs of electrons of the oxygen atom to the aromatic ring (electron-donating) [Schroeter et al., 1998; Amunugama and Rodgers, 2002c]. In other words, the rather weak π-electron donating property of the hydroxy substituent should not lead to drastic increase in absolute Ag⁺ binding energies. From another point of view, both the molecular polarizabilities and quadrupole moments of X-PhOH are, on average, only slightly greater by ~8 % than X-Bz (Table 4.3 and 7.4).

<table>
<thead>
<tr>
<th></th>
<th>Q_{zz}^a</th>
<th>α^b</th>
<th>μ^c</th>
</tr>
</thead>
<tbody>
<tr>
<td>PhOH</td>
<td>-26.3</td>
<td>9.0 (11.1)</td>
<td>1.59 (1.224)</td>
</tr>
<tr>
<td>Me-PhOH</td>
<td>-25.0</td>
<td>10.9</td>
<td>1.51 (1.54)</td>
</tr>
<tr>
<td>Et-PhOH</td>
<td>-24.5</td>
<td>12.6</td>
<td>1.53</td>
</tr>
<tr>
<td>i-Pr-PhOH</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>i-Bu-PhOH</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>MeO-PhOH</td>
<td>-29.3</td>
<td>11.7</td>
<td>2.96</td>
</tr>
<tr>
<td>EtO-PhOH</td>
<td>-28.0</td>
<td>13.5</td>
<td>2.84</td>
</tr>
<tr>
<td>NO₂-PhOH</td>
<td>-0.2</td>
<td>12.3</td>
<td>4.21</td>
</tr>
</tbody>
</table>

^a Theoretical Quadrupole moment, Q_{zz}, calculated at the B3LYP/6-31G(d,p)/MP2/3-21G(d) level of theory, in units of 10^{-40} C m^2.

^b Theoretical molecular polarizability, α, calculated at the B3LYP/6-31G(d,p)/MP2/3-21G(d) level of theory, in units of 10^{24} cm^3. Experimental molecular polarizabilities taken from [Lide, 2002-2003] are shown in parenthesis.

^c Theoretical molecular dipole moments, μ, calculated at the MP2/3-21G(d) level of theory, in units of debye D. Experimental molecular dipole moment taken from [Lide, 2002-2003] is shown in parenthesis.
Figure 7.7 Direction of molecular dipole moments in (a) 4-methoxyphenol (MeO-PhOH) and (b) 4-nitrophenol (NO$_2$-PhOH).

7.2.4 Competition between Cation-π and Non-π Interactions: Ag (I) Cation Bound Complexes of 4-Nitrophenol

As in the case of X-Indole, Ag$^+$ prefers to bind to non-π binding sites in 4-nitrophenol (NO$_2$-PhOH) and in the direction of dipole moment (Figures 7.6 and 7.7b). Thus, the situation discussed in the previous chapter should also be applicable to NO$_2$-PhOH, i.e. ion-dipole and ion-induced dipole interactions are mainly responsible for Ag$^+$ binding to the −NO$_2$ substituent.

7.2.5 Comparison of Ag (I) Cation Affinities with Other Metal Cation Affinities of Phenol

To our knowledge, there is no literature report on the cation binding to substituted phenols (X-PhOH). However, a few theoretical studies on metal cation binding with phenol (PhOH) has been reported, and the findings are summarized in Table 7.5. As discussed in Section 7.2.2, theoretical studies indicate that Ag$^+$ binds more strongly to the cation-π binding site of phenol than the non-π hydroxyl-O binding site by 21.6
kJ mol\(^{-1}\). Similar to the case of Ag\(^{+}\), Amunugama and Rodgers found that the theoretical Li\(^{+}\) and K\(^{+}\) bound cation-\(\pi\) complexes of phenol, determined at the MP2(full)/6-311+G(2d,2p)// B3LYP/6-31G* level of theory, are more energetically favored over the corresponding non-\(\pi\) conformers (i.e. the hydroxyl-O binding site) by 7.5 and 1.5 kJ mol\(^{-1}\), respectively [Amunugama and Rodgers, 2002c]. In contrast, the Na\(^{+}\) and Rb\(^{+}\) non-\(\pi\) binding modes are found to be more stable than the cation-\(\pi\) binding modes by 0.6 and 1.5 kJmol\(^{-1}\), respectively. The behaviour of Cs\(^{+}\) is intermediate between Na\(^{+}\) and Rb\(^{+}\). However, with neglect of ZPE and BSSE corrections, the cation-\(\pi\) complexes of Li\(^{+}\), Na\(^{+}\), K\(^{+}\), Rb\(^{+}\) and Cs\(^{+}\) were found to be more stable than the corresponding non-\(\pi\) conformers by 13.1, 2.8, 3.3, 0.4 and 0.7 kJ mol\(^{-1}\), respectively. This shows that the alkali metal cation affinities of phenol for the cation-\(\pi\) and non-\(\pi\) binding modes are close to each other and so the experimental threshold-CID binding energies could be ascribed to any one or a mixture of the cation-\(\pi\) and non-\(\pi\) binding complexes. Our results also revealed a similar situation for the Ag\(^{+}\)-phenol interaction. Similar to the other aromatic systems reported, the Ag\(^{+}\) binding affinity of phenol is larger than Na\(^{+}\) and comparable to Li\(^{+}\), which is due to the presence of covalent interaction as explained in the previous chapters.

Likewise, Ryzhov and Dunbar predicted by B3LYP-density functional (DFT) calculations that the two binding sites (cation-\(\pi\) and non-\(\pi\)) of phenol are close in energy for Na\(^{+}\), Mg\(^{+}\) and Al\(^{+}\), but the cation-\(\pi\) complexes show greater stability over the non-\(\pi\) conformers for Cr\(^{+}\) and Fe\(^{+}\) [Ryzhov and Dunbar, 1999]. The B3LYP-DFT results are consistently smaller than the experimental radiative kinetics value by 12 – 42 kJ mol\(^{-1}\), the “best-estimate” values for Na\(^{+}\) and K\(^{+}\) are ~ 15% greater than
the threshold-CID values measured by Amunugama and Rodgers [Amunugama and Rodgers, 2002c].

<table>
<thead>
<tr>
<th>M⁺</th>
<th>ΔH₀ (Expt)</th>
<th>ΔH₀ (Theory)</th>
<th>Best-estimate ΔH°</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag⁺</td>
<td>158.3⁻</td>
<td>159.9⁻</td>
<td>–</td>
</tr>
<tr>
<td>Li⁺</td>
<td>178.5 ± 16.1ᵇ</td>
<td>156.8ᵇ</td>
<td>–</td>
</tr>
<tr>
<td>Na⁺</td>
<td>102.3 ± 3.4ᵇ</td>
<td>102.6ᵇ</td>
<td>117.2 ± 12.6</td>
</tr>
<tr>
<td></td>
<td>98.5 ± 3.4ᶜ</td>
<td>91.1ᶜ</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>104.2ᶠ</td>
<td></td>
</tr>
<tr>
<td>K⁺</td>
<td>74.0 ± 3.4ᵇ</td>
<td>79.2ᵇ</td>
<td>83.7 ± 12.6</td>
</tr>
<tr>
<td>Rb⁺</td>
<td>68.7 ± 4.4ᵇ</td>
<td>65.2ᵇ</td>
<td>–</td>
</tr>
<tr>
<td>Cs⁺</td>
<td>65.3 ± 4.9ᵇ</td>
<td>58.4ᵇ</td>
<td>–</td>
</tr>
<tr>
<td>Mg⁺</td>
<td>159.8ᵈ</td>
<td>137.7ᶠ</td>
<td>154.8 ± 16.7</td>
</tr>
<tr>
<td>Al⁺</td>
<td>156.9ᵈ</td>
<td>134.3ᶠ</td>
<td>159.0 ± 16.7</td>
</tr>
<tr>
<td>Cr⁺</td>
<td>171.5 ± 10.5ᵉ</td>
<td>164.0ᶠ</td>
<td>179.9 ± 16.7</td>
</tr>
<tr>
<td></td>
<td>206.3ᵈ</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fe⁺</td>
<td>209.2 ± 10.5ᵉ</td>
<td>276.1ᶠ</td>
<td>251.0 ± 25.1</td>
</tr>
<tr>
<td></td>
<td>&gt; 238.5ᵈ</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Co⁺</td>
<td>255.2 ± 11.3ᵉ</td>
<td>–</td>
<td>–</td>
</tr>
</tbody>
</table>

⁻ Our work as shown in Table 5.3.
ᵇ Threshold-CID ΔH₀ and calculations at the MP2(full)/6-311+G(2d,2p)// B3LYP/6-31G* level with neglect of ZPE and BSSE corrections, ΔH₀. [Amunugama and Rodgers, 2000c]
ᶜ Threshold-CID ΔH₀ and calculations at the MP2(full)/6-311+G(2d,2p)// B3LYP/6-31G* level of theory, ΔH₀. [Armentrout and Rodgers, 2000]
ᵈ Radiative association kinetic analysis, [Ryzhov and Dunbar, 1999].
ᵉ Kinetic method, [Schroeter et al., 1998]
ᶠ DFT calculation at B3LYP level of theory with neglect of ZPE and BSSE corrections, ΔH₀, [Ryzhov and Dunbar, 1999].
ᵉ ‘Best estimates’ of absolute binding energy based on available experimental evidence supplemented by quantum chemistry comparative estimates, [Ryzhov and Dunbar, 1999].
Schroeter and co-workers also determined the binding energies of Cr\(^+\), Fe\(^+\) and Co\(^+\) to PhOH by the kinetic method measurements [Schroeter et al., 1998]. Their affinity value for Cr\(^+\) is in agreement with that measured by Ryzhov and Dunbar, but the Fe\(^+\) affinity value is significantly below the "best-estimate" value and even the lower limit reported by Ryzhov and Dunbar [Ryzhov and Dunbar, 1999]. Amunugama and Rodgers suggested that the Fe\(^+\)-PhOH complex accessed in the two laboratories could be in different electronic states [Amunugama and Rodgers, 2002c].

Despite the uncertainties in experimental measurements, the trend of transition metal cation affinities of Ag\(^+\) ~ Cr\(^+\) < Fe\(^+\) < Co\(^+\) for phenol is observed, similar to that found in alkylbenzenes (Chapter 4). The observed trend is rationalized in terms of different extents of charge-transfer interaction present in the cation-\(\pi\) binding mode.

The binding energies of Mg\(^+\) and Al\(^+\) to PhOH are much stronger than those of alkali metal cations because of greater charge transfer contribution in Mg\(^+\) and Al\(^+\) (via the sp-polarization effect, which have been mentioned earlier) [Rodger et al., 2000; Amunugama and Rodgers, 2002c; 2003a]. The binding energies of Mg\(^+\) and Al\(^+\) are similar to that of Ag\(^+\), further confirming that charge-transfer interaction is present and is important in determining the metal cation affinities relative to those of alkali metal cations, which are predominantly electrostatic in nature.
Chapter 8  Conclusion

8.1  Summary of Major Findings

In this study, the Ag$^+$ affinities of model aromatic-π systems in gas phase, including alkylbenzenes, substituted naphthalenes, substituted indoles and substituted phenols, were measured by the mass spectrometric kinetic method. Our approach started with experimental validation of the theoretical Ag$^+$ affinity values of alkylbenzenes. The experimental Ag$^+$ affinities of other model aromatic-π ligands were then determined by using alkylbenzenes as reference compounds, and further validated by measurement of and comparison with known Ag$^+$ affinities reported in the literature.

The experimental Ag$^+$ affinities measured in this study are in very good agreement with theoretical ab initio values estimated at the CCSD(T)/[HW(f), 6-31G+(d)]/MP2/[HW, 3-21G(d)] level; the difference between experimental and theoretical values are generally within the limits of uncertainty of the experimental measurements (± 10 – 14 kJ mol$^{-1}$).

The experimental Ag$^+$ affinities at 0K of 15 alkylbenzenes determined by the standard kinetic method (as shown in parenthesis in kJ mol$^{-1}$) are: Bz (157.1) < Tol (168.3) < EtBz (175.9) < m-Xy (180.6) < n-PrBz (182.2) < i-PrBz (182.6) < n-BuBz (186.4) < i-BuBz (187.1) < s-BuBz (188.2) < t-BuBz (189.0) < 1,3,5-Me$_3$Bz (192.6) < n-PnBz (193.7) < 1,2,4,5-Me$_4$Bz (200.3) < Me$_2$Bz (209.1) < Me$_6$Bz (216.5).

The experimental Ag$^+$ affinities at 0K of 9 substituted naphthalenes are: Nap (176.8) < HO-Nap (178.1) < Me-Nap (188.5) < MeO-Nap (191.0) < Et-Nap (193.9) ~ CN-Nap (194.8) < EtO-Nap (195.3) < i-Pr-Nap (198.0) < NH$_2$Nap (211.1).
The experimental Ag⁺ affinities at 0K of 10 substituted indoles determined by the standard kinetic method are: 5-HO-Indole (200.7) < Indole (207.3) < 5-MeO-Indole (213.2) ~ 3-Me-Indole (214.2) ~ 5-Me-Indole (215.1) ~ 6-Me-Indole (216.7) < 7-Me-Indole (219.4) < 1-Me-Indole (221.9) ~ 2-Me-Indole (223.4) ~ 7-MeO-Indole (224.9). The Ag⁺ affinities of two substituted indoles determined by the extended kinetic method are: 5-NO₂-Indole (196.6) and 5-CN-Indole (213.7).

The experimental Ag⁺ affinities at 0K of 7 substituted phenols are: PhOH (158.3) < Me-PhOH (165.2) < Et-PhOH (172.2) < i-Pr-PhOH (178.8) < t-Bu-PhOH (183.7) and MeO-PhOH (165.6) < EtO-PhOH (170.9). The Ag⁺ affinity of nitrophenol determined by the extended kinetic method is: 166.5 kJmol⁻¹.

Our experimental and theoretical studies show that Ag⁺ cation-π binding is the most stable and preferred binding mode for alkyl, hydroxyl and alkoxy substituted benzenes, naphthalenes, indoles and phenols, indicating that non-covalent cation-π interaction is indeed strong enough to compete against Ag⁺ binding to oxygen-donor binding sites. On the other hand, Ag⁺ cation-π binding is not as strong as Ag⁺ binding to heteroatom nitrogen/oxygen binding sites min –CN, -NH₂ and –NO₂ substituted benzenes, naphthalens, indoles and phenols.

Our results suggested that Ag⁺ cation-π interaction is mainly electrostatic in nature, though charge-transfer covalent interaction is noticeably present. The relative stability of the cation-π and non-π binding to O/N heteroatom sites is mainly determined by the interplay of electrostatic ion-quadrupole, ion-dipole, ion-induced dipole interactions, as well as the extent of charge-transfer (covalent) interaction present in the Ag⁺-ligand complexes. Ag⁺ binding to substituted naphthalenes and substituted indoles are measured to be ~ 9 – 12 % and ~ 11 – 33 % stronger than of
substituted benzenes, respectively, indicating that both the quadrupole moment and polarizability of fused aromatic ring systems could greatly enhance cation-π binding energy. Substituted phenols show comparable Ag⁺ affinities as substituted benzenes, i.e., the phenolic-OH group does not have any significant effect on Ag⁺-cation-π binding affinities. The increasing trend of Ag⁺ affinities within a series of substituted benzenes naphthalenes, indoles and phenols is shown to be mainly due to the increasing molecular polarizability of the substituted aromatic systems, which could enhance the Ag⁺ affinities via both the ion-induced dipole and charge-transfer interactions.

Ag⁺-cation-π binding affinities are comparable to that of Li⁺, but greater than Na⁺ affinities, even though the ionic radius of Ag⁺ is larger than that of Na⁺. The greater Ag⁺ affinity is attributed to the noticeable presence of charge-transfer (covalent) interactions, i.e., π → Ag⁺ (σ – donor interaction) and Ag⁺ → π* (4d¹⁰ back donation), present in Ag⁺-ligand interactions. In fact, the order of transitional metal cation-π affinities: Co⁺ > Fe⁺ > Cr⁺ ~ Ag⁺ can be rationalized in terms of the variation in charge-transfer interactions for different transition metal cations.
8.2 Suggestion for Further Work

In the present study, the relative importance of ion-quadrupole, ion-dipole, ion-induced dipole and charge-transfer interactions, Ag\(^+\)-cation-\(\pi\) and non-\(\pi\) modes of binding are deduced by correlation to molecular quadrupole moment, dipole moment and polarizability of the ligands. This approach could only explain qualitatively the Ag\(^+\) affinity trends in homologous series of substituted benzenes, naphthalenes, indoles and phenols.

A quantitative approach can be carried out by Natural Bond Order (NBO) analysis, from which the relative contributions of charge-transfer and electrostatic interactions can be elucidated. Such studies would better reveal the origin of the Ag\(^+\) affinity interactions, and is important in furthering the understanding and application of cation-\(\pi\) interactions in biological and material science.
Appendix I: Figures and Tables for Chapter 4

<table>
<thead>
<tr>
<th>Alkylbenzene, RBz</th>
<th>ln([Ag + RBz])\textsuperscript{+}/[Ag + RBz])</th>
<th>ln([Ag + RBz])\textsuperscript{+}/[Ag + Bz]\textsuperscript{+})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bz</td>
<td>0.00 ± 0.00</td>
<td></td>
</tr>
<tr>
<td>Tol</td>
<td>0.97 ± 0.06</td>
<td>0.97 ± 0.06</td>
</tr>
<tr>
<td>EtBz</td>
<td>0.96 (0.31)</td>
<td>1.63 ± 0.07</td>
</tr>
<tr>
<td>m-Xy</td>
<td>1.14</td>
<td>1.94 ± 0.10</td>
</tr>
<tr>
<td>n-PrBz</td>
<td>0.57 (0.04)</td>
<td>2.16 ± 0.10</td>
</tr>
<tr>
<td>i-PrBz</td>
<td>0.25 (0.04)</td>
<td>2.20 ± 0.11</td>
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<td>n-BuBz</td>
<td>0.37 (0.08)</td>
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<td>i-BuBz</td>
<td>0.36 (0.09)</td>
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<td>s-BuBz</td>
<td>0.49 (0.08)</td>
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<td>t-BuBz</td>
<td>0.18 (0.02)</td>
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<tr>
<td>1,3,5-Me\textsubscript{3}Bz</td>
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<td>2.87 ± 0.15</td>
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<tr>
<td>n-PnBz</td>
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<td>3.11 ± 0.15</td>
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<tr>
<td>1,2,4,5-Me\textsubscript{4}Bz</td>
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<td>3.61 ± 0.15</td>
</tr>
<tr>
<td>Me\textsubscript{2}Bz</td>
<td>1.30</td>
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</tr>
<tr>
<td>Me\textsubscript{3}Bz</td>
<td>0.54</td>
<td>4.88 ± 0.16</td>
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Figure A4.1  Experimentally measured \(\ln([\text{Ag} + \text{RBz}]^\text{+})/ [\text{Ag} + \text{RBz}]^\text{+} \) values for the high-energy (4.7 keV, laboratory frame) He-CID of Ag\textsuperscript{+} bound heterodimers of alkylbenzenes (T\textsubscript{eff} = 1,460 K). The \(\ln([\text{Ag} + \text{RBz}]^\text{+})/ [\text{Ag} + \text{RBz}]^\text{+} \) values are the logarithm of ion intensity ratios. The data presented under the heading \(\ln([\text{Ag} + \text{RBz}]^\text{+})/ [\text{Ag} + \text{Bz}]^\text{+} \) are average ± standard deviation of cumulative values expressed relative to benzene (Bz). The values in ( ) are the arithmetic difference of measured values having a common reference alkylbenzene.
Figure A4.2  Calibration plot of cumulative $\ln([\text{Ag} + \text{RBz}]^+ / [\text{Ag} + \text{Bz}]^+)$ values obtained under high-energy He-CID (Figure A4.1, $T_{\text{eff}} = 1,460$ K) conditions versus theoretical Ag$^+$ affinities of alkylbenzenes of Tol, $m$-Xy, 1,3,5-Me$_3$Bz, Me$_6$Bz (●) at 0K. The absolute Ag$^+$ affinities of other alkylbenzenes (○) were determined by interpolation and extrapolation of the plot.
Figure A4.3  (a) Plot of $\ln([\text{Ag} + \text{(Acetone)}]^+ / [\text{Ag} + \text{RBz}]^+)$ versus $[\Delta H_{[\text{Ag} + \text{RBz}]^+} - \Delta H_{\text{avg}}]$ at different collision energies and  (b) plot of $[\Delta G_{\text{app}}[^{\text{Ag} + \text{(Acetone)}}]^+] / RT_{eff}$ versus $1/RT_{eff}$ for the heterodimers $[(\text{Acetone}) + \text{Ag} + \text{RBz}^+(\text{RBz} = \text{Tol, EtBz, m-Xy, i-PrBz, t-BuBz})]$. The Ag$^+$ affinity of acetone, $\Delta H_{[\text{Ag} + \text{(Acetone)}]^+}$, and the $\Delta(\Delta S_{\text{Ag}})^{\text{app}}$ term are expressed as un-weighted values.
<table>
<thead>
<tr>
<th>RBz</th>
<th>$\Delta H_{[\text{Ag} + \text{RBz}]}^{a}$</th>
<th>$\Delta H_{[\text{H} + \text{RBz}]}^{b}$</th>
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<td>791.9</td>
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</tr>
<tr>
<td>s-BuBz</td>
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<td>-</td>
</tr>
<tr>
<td>t-BuBz</td>
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<td>-</td>
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<tr>
<td>Me$_6$Bz</td>
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<td>860.6</td>
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* $^{a}$ Experimental Ag$^+$ affinities, $\Delta H_{\text{Ag}}$, as shown in Table 4.1.

* $^{b}$ Experimental proton affinities taken from [Hunter and Lias, 1998].
Appendix II: Figures and Tables for Chapter 5

(a)

<table>
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<th>Ligands, L</th>
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<th>$\ln([\text{Ag} + L]^+/[\text{Ag} + \text{Tol}]^+)$</th>
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<td>Tol</td>
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<td>m-Xy</td>
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<td>Me-Nap</td>
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<td>1.61 ± 0.07</td>
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<tr>
<td>1,3,5-Me$_3$Bz</td>
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<td>0.37 0.26</td>
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<td>0.40 0.13</td>
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<td></td>
<td></td>
<td>3.30 ± 0.10</td>
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</table>

(b)

<table>
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<th>Ligands, L</th>
<th>$\ln([\text{Ag} + L_1]^+/[\text{Ag} + L_2]^+)$</th>
<th>$\ln([\text{Ag} + L]^+/[\text{Ag} + \text{Tol}]^+)$</th>
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<tr>
<td>Tol</td>
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</tr>
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<td>HO-Nap</td>
<td>1.86</td>
<td>1.02 0.84</td>
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<tr>
<td></td>
<td></td>
<td>0.84 ± 0.00</td>
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<tr>
<td>m-Xy</td>
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<td>0.99 0.24</td>
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<tr>
<td>MeO-Nap</td>
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<td>1.08 ± 0.03</td>
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<tr>
<td>1,3,5-Me$_3$Bz</td>
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<tr>
<td>EtO-Nap</td>
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<td></td>
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<td>2.21 ± 0.07</td>
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Figure A5.1 Experimentally measured $\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ values for the low-energy (25 eV, laboratory frame) Ar-CID of Ag$^+$ bound heterodimers of substituted naphthalenes (T$_{\text{eff}}$ = (a) 1,563 K and (b) 1,454 K). The $\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ values are the logarithm of ion intensity ratios. The data presented under the heading The $\ln([\text{Ag} + L]^+ / [\text{Ag} + \text{Tol}]^+)$ are average ± standard deviation of cumulative values expressed relative to toluene (Tol).
Figure A5.2  Calibration plots of cumulative $ln([Ag + L]^+ / [Ag + Tol]^+)$ values (Figure A5.1) obtained under low-energy (25eV, laboratory frame) Ar-CID ($T_{eff} = (a) 1,563K$ and (b) $1,454K$) conditions versus Ag$^+$ affinities of reference alkylbenzenes of Tol, m-Xy, 1,3,5-Me$_3$Bz (●) at 0K. The absolute Ag$^+$ affinities of substituted naphthalenes were determined by interpolation and extrapolation of the plot.
Figure A5.3  Triple quadrupole MS/MS spectrum of the $[(\text{NH}_2\text{-Nap}) + \text{Ag} + (1,2,4,5\text{-Me}_4\text{Bz})]^+$ heterodimer using argon as collision gas at a various collision energies (laboratory frame) $(\text{NH}_2\text{-Nap} = 2\text{-aminonaphthalene}$ and $1,2,4,5\text{-Me}_4\text{Bz} = 1,2,4,5\text{-tetramethylbenzene})$. 

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Figure A5.4  (a) Plot of \( \ln(\frac{[\text{Ag} + (\text{NH}_2\text{Nap})]^+}{[\text{Ag} + \text{RBz}]^+}) \) versus \( \Delta H_{\text{Ag} + 2\text{RBz}}^+ - \Delta H_{\text{Ag}} \) at different collision energies and (b) plot of \( \frac{\Delta G_{\text{Ag} + (\text{NH}_2\text{Nap})]^+}}{\Delta H_{\text{Ag}}} / \frac{\Delta H_{\text{Ag}}}{\Delta H_{\text{Ag}}} \) versus \( 1/RT_{\text{eff}} \) for the heterodimers \( [(\text{CN-Nap}) + \text{Ag} + \text{RBz}]^+ \) (\( \text{NH}_2\text{Nap} = 2\)-aminonaphthalene and \( \text{RBz} = m\)-\( \text{Xy}, 1,3,5\)-\( \text{Me}_3\text{Bz}, 1,2,4,5\)-\( \text{Me}_4\text{Bz}, \text{Me}_2\text{Bz} \)). The \( \text{Ag}^+ \) affinity of \( \text{NH}_2\text{Nap}, \Delta H_{\text{Ag} + (\text{NH}_2\text{Nap})]^+}, \) and the entropy changes, \( \Delta(\Delta S_{\text{Ag}^+})^{\text{app}} \), are expressed as un-weighted values.
Table A5.1  Experimental Ag⁺ affinities, $\Delta H_0$, at 0K (in kJ mol⁻¹) of substituted benzenes (X-Bz) and substituted naphththalenes (X-Nap)

<table>
<thead>
<tr>
<th>Substituent (X)</th>
<th>$\Delta H_0$ (Exp) X-Bz</th>
<th>$\Delta H_0$ (Exp) X-Nap</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>157.1⁺</td>
<td>176.8</td>
</tr>
<tr>
<td>CH₃</td>
<td>168.3⁺</td>
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<tr>
<td>C₂H₅</td>
<td>175.9⁺</td>
<td>193.9</td>
</tr>
<tr>
<td>i-C₃H₇</td>
<td>182.6⁺</td>
<td>198.0</td>
</tr>
<tr>
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<tr>
<td>OC₂H₅</td>
<td>178.6ᵇ</td>
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</tr>
<tr>
<td>NO₂</td>
<td>150.9ᵇ</td>
<td>–</td>
</tr>
<tr>
<td>CN</td>
<td>–</td>
<td>194.8</td>
</tr>
<tr>
<td>NH₂</td>
<td>194.2ᵇ</td>
<td>211.1</td>
</tr>
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</table>

⁻ Experimental Ag⁺ affinities, $\Delta H_0$, measured by standard kinetic method in Table 4.1.

ᵇ Experimental Ag⁺ affinities, $\Delta H_0$, measured by kinetic method measurements by Miss S. K. Wo of our research group.

ᶜ Experimental Ag⁺ affinities, $\Delta H_0$, measured by standard kinetic method as shown in Table 5.1.
Appendix III: Figures and Tables for Chapter 6

<table>
<thead>
<tr>
<th>Ligands, L</th>
<th>$\ln([Ag + L_1]^+/[Ag + L_2]^+)$</th>
<th>$\ln([Ag + L]^+/[Ag + Bz]^+)$</th>
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<tbody>
<tr>
<td>Bz</td>
<td>0.00 ± 0.00</td>
<td>1.33 ± 0.00</td>
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<td>Tol</td>
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</tr>
<tr>
<td>m-Xy</td>
<td>2.48</td>
<td>2.51 ± 0.06</td>
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<td>1,3,5-Me_3Bz</td>
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<td>3.65 ± 0.07</td>
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<tr>
<td>5-HO-Indole</td>
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<tr>
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<td>3-Me-Indole</td>
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<td>7-MeO-Indole</td>
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<td>6.68 ± 0.11</td>
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Figure A6.1  Experimentally measured $\ln([Ag + L_1]^+ / [Ag + L_2]^+)$ values for the low-energy (40 eV, laboratory frame) Ar-CID of Ag$^+$ bound heterodimers of substituted indoles ($T_{eff} = 1,637$ K). The $\ln([Ag + L]^+ / [Ag + Bz]^+)$ values are the logarithm of ion intensity ratios. The data presented under the heading $\ln([Ag + L]^+ / [Ag + Bz]^+)$ are average ± standard deviation of cumulative values expressed relative to benzene (Bz). The values in ( ) are the arithmetic difference of measured values having a common reference ligands.
<table>
<thead>
<tr>
<th>Ligands L</th>
<th>$\ln([\text{Ag} + L_1]^+/[\text{Ag} + L_2]^+)$</th>
<th>$\ln([\text{Ag} + L]^+/[\text{Ag} + (5\text{-HO-Indole})]^+)$</th>
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<td>Indole</td>
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<td>0.14 ± 0.01</td>
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<td>3-Me-Indole</td>
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<td>5-Me-Indole</td>
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<td>1.46 ± 0.03</td>
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<td>0.50</td>
<td>0.36</td>
</tr>
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<td></td>
<td>0.36</td>
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**Figure A6.2** Experimentally measured $\ln([\text{Ag} + L_1]^+/ [\text{Ag} + L_2]^+)$ values for the high-energy (4.7 keV, laboratory frame) He-CID of Ag$^+$ bound heterodimers of substituted indoles. The $\ln([\text{Ag} + L_1]^+/ [\text{Ag} + L_2]^+)$ values are the logarithm of ion intensity ratios. The data presented under the heading The $\ln([\text{Ag} + L]^+/ [\text{Ag} + (5\text{-HO-Indole})]^+)$ are average ± standard deviation of cumulative values expressed relative to 5-hydroxyindole (5-HO-Indole). The values in () are the arithmetic difference of measured values having a common reference ligands.
Figure A6.3 Low-energy ((a) 20eV and (b) 40eV, laboratory frame) Ar-CID mass spectra of Ag (I) cation bound heterodimer, [Indole + Ag + (m-Xy)]^+, (m-Xy = m-xylene).
Figure A6.4  Calibration plot of cumulative \( \ln([\text{Ag} + \text{L}]^+ / [\text{Ag} + \text{Bz}]^+) \) values (Figure A6.1) obtained under low-energy (40eV \( T_{\text{eff}} = 1,637 \) K), laboratory frame) Ar-CID conditions versus Ag\(^+\) affinities of reference alkylbenzenes of Bz, Tol, m-Xy, 1,3,5-Me\(_3\)Bz (●) at 0K. The absolute Ag\(^+\) affinities of alkyl and alkoxy substituted indoles were determined by interpolation and extrapolation of the plot.
Figure A6.5  Triple quadrupole MS/MS spectrum of the [(5-CN-Indole) + Ag + Me₆Bz]⁺ heterodimer using argon as collision gas at a various collision energies (laboratory frame) (5-CN-Indole = 5-cyanoindole and Me₆Bz = hexamethylbenzene).
Figure A6.6  (a) Plot of \( \ln([\text{Ag} + (5\text{-CN-Indole})]^+) / [\text{Ag} + \text{RBz}^+] \) versus \( \Delta H_{[\text{Ag} + \text{RBz}^+] - \Delta H_{\text{Avg}}} \) at different collision energies and (b) plot of \( \frac{[\Delta G_{\text{app}}^{\text{app}}_{[\text{Ag} + (5\text{-CN-Indole})]^+} - \Delta H_{\text{Avg}}]}{\text{RT}_{\text{eff}}} \) versus \( 1/\text{RT}_{\text{eff}} \) for the heterodimers [(5-CN-Indole) + Ag + RBz]^+ (5-CN-Indole = 5-cyanoindole and RBz = 1,3,5-Me,Bz, 1,2,4,5-Me,Bz, Me,Bz, Me,Bz). The Ag⁺ affinity of 5-CN-Indole, \( \Delta H_{[\text{Ag} + (5\text{-CN-Indole})]^+} \), and the entropy changes, \( \Delta (\Delta S_{\text{Ag}}^{\text{app}}) \), are expressed as un-weighted values.
Figure A6.7  Binding sites and geometries of stable conformers Ag⁺-(5-HO-Indole) complexes at MP2/[HW, 3-21G(d)] level of theory.
Figure A6.8  Binding sites and geometries of stable conformers Ag⁺-(7-MeO-Indole) complexes at MP2/[HW, 3-21G(d)] level of theory.
Figure A6.9  Binding sites and geometries of stable conformers Ag⁺-(5-NO₂-Indole) complexes at MP2/[HW, 3-21G(d)] level of theory.
## Appendix IV: Figures and Tables for Chapter 7

(a)

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<tr>
<th>Ligands, L</th>
<th>$\ln([\text{Ag} + L_1]^+/[\text{Ag} + L_2]^+)$</th>
<th>$\ln([\text{Ag} + L]^+/[\text{Ag} + \text{Bz}]^+)$</th>
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<tbody>
<tr>
<td>Bz</td>
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<td>0.00 ± 0.00</td>
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<td>2.25 ± 0.10</td>
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(b)

<table>
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<tr>
<th>Ligands, L</th>
<th>$\ln([\text{Ag} + L_1]^+/[\text{Ag} + L_2]^+)$</th>
<th>$\ln([\text{Ag} + L]^+/[\text{Ag} + \text{Bz}]^+)$</th>
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<td>Bz</td>
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<td>1.82 ± 0.12</td>
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**Figure A7.1** Experimentally measured $\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ values for the low-energy (40 eV, laboratory frame) Ar-CID of Ag$^+$ bound heterodimers of (a) alkyl substituted phenols ($T_{\text{eff}} = 1,448$ K) (b) alkoxy substituted phenols ($T_{\text{eff}} = 1,552$ K). The $\ln([\text{Ag} + L_1]^+ / [\text{Ag} + L_2]^+)$ values are the logarithm of ion intensity ratios. The data presented under the heading $\ln([\text{Ag} + L]^+ / [\text{Ag} + \text{Bz}]^+)$ are average ± standard deviation of cumulative values expressed relative to benzene (Bz).
Figure A7.2 Low-energy (a) 25eV and (b) 40eV, laboratory frame) Ar-CID mass spectra of Ag (I) Cation bound heterodimer, [(MeO-PhOH) + Ag + Bz]^+, (MeO-PhOH = 4-methoxyphenol and Bz = benzene).
Figure A7.3  Calibration plots of cumulative $\ln([\text{Ag} + \text{L}]^+ / [\text{Ag} + \text{Bz}]^+)$ values (Figure A7.1) obtained under low-energy (40eV, laboratory frame) Ar-CID ($T_{\text{eff}} =$ (a) 1,448 K and (b) 1,552 K) conditions versus Ag$^+$ affinities of reference alkylbenzenes of Bz, Tol, m-Xy (●) at 0K. The absolute Ag$^+$ affinities of alkyl and alkoxy substituted phenols (○) were determined by interpolation and extrapolation of the plot.
Figure A7.4 Binding sites and geometries of stable conformers Ag⁺-(MeO-PhOH) complexes optimized at MP2/[HW, 3-21G(d)] level of theory.
Figure A7.5  Binding sites and geometries of stable conformers $\text{Ag}^+-(\text{NO}_2^-\cdot\text{PhOH})$ complexes optimized at MP2/[HW, 3-21G(d)] level of theory.
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II. Conference Presentations


