(2989 words)

An investigation on the impact of electronegativity on experimental and theoretical lattice energies

Introduction

Lattice enthalpy is the energy absorbed when one mole of a solid ionic compound is separated into gaseous ions under standard conditions (Brown, 2014).

$$MX(s) \rightleftharpoons M^{+}(g) + X^{-}(g)$$

Theoretical lattice enthalpies: Are calculated assuming the ionic crystal is made up from perfectly spherical ions with only interaction of electrostatic forces between them. The energy depends on the product of the ionic charges and the sum of the ionic radii of the ionic compound:

- *An increase in the ionic radius of one of the ions decreases the attraction between the ions*
- *An increase in the ionic charge increases the ionic attraction between the ions*[1]

Theoretical lattice enthalpy for one mole is calculated with the following equation;

$$\Delta H\,^{0}_{lat} = \frac{Kum}{R_{M^{m+}} + R_{X^{m-}}}$$

This is because the interactions between ions have to be considered as a solid crystal form. Thus, the "attraction between the positive and negative ions predominates over the repulsion of ions with the same charge as ions are generally surrounded by neighbouring ions of opposite charge" [2]

K is a constant that depends on the geometry of the lattice and n and m are the magnitude of charges on the ions. As the ionic radii can be determined from X-ray diffraction measurements of the crystal, theoretical values can be calculated once the geometry of the solid lattice is known"[3].

However, I have decided to use the Kapustinskii equation and modified Born-Lande equation to calculate the theoretical lattice enthalpy values because I found it difficult to measure the ion interactions of solid crystal forms to use the previous equation.

I used the Kapustinskii equation and modified Born-Lande equation to calculate the theoretical lattice enthalpy values.

Kapustinskii equation is defined below:[4]

$$\Delta H^{0}_{\text{lattice}} = K \cdot \frac{\nu \cdot |z^{+}| \cdot |z^{-}|}{r^{+} + r^{-}} \cdot \left(1 - \frac{d}{r^{+} + r^{-}}\right) \times 10^{-3}$$

K = 1.20200 × 10^{-4} J m mol^{-1}
d = 3.45 × 10^{-11} m
$r^{+} + r^{-}$ = sum of ionic radii (m)
ν = number of ions in empirical formula
z^{+}, z^{-} = numeric charge of ion

Born-Lande equation is defined below:[5]

---

[1] (Brown, 2014)
[2] (Brown, 2014).
[3] (Brown, 2014).
[4] (Kapustinskii Equation, 2020
[5] (Born-Lande Equation, 2020)

$$\Delta H_{lattice} = -\frac{N_A M z^+ z^- e^2}{4\pi\varepsilon_0 r_0}\left(1 - \frac{1}{n}\right)$$

  - $N_A = Avogadro\ constant$
  - $M = Madelung\ constant, relating\ to\ the\ geometry\ of\ the\ crystal$
  - $z^+, z^- = numeric\ charge\ number\ of\ ion$

- $e = elementary\ charge, 1.6022 \times 10^{-19} C$
- $\varepsilon_0 = permittivity\ of\ free\ space, 8.854 \times 10^{-12} C^2 J^{-1} m^{-1}$
- $r_0 = distance\ between\ ions$
- $n = Born\ exponent$ (determined experimentally by measuring the compressibility of the solid, or derived theoretically)

Experimental lattice enthalpy

A Born-Haber cycle is used because this value can't be directly measured:

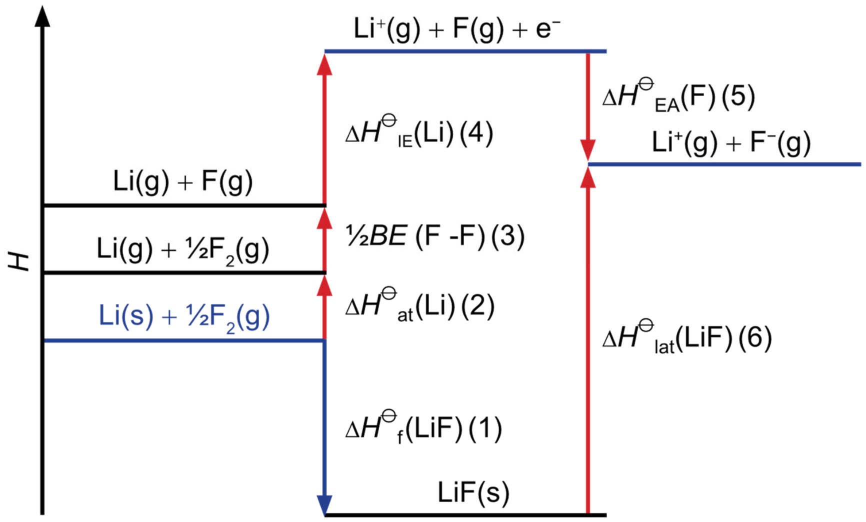


Figure 1 - Born-Haber cycle for UF

Born-Haber cycle can be written in energy level diagrams (Figure 1)[6].

The Born-Haber cycle for lithium fluoride is constructed by:

1. Enthalpy change of formation: formation of elements in the standard state from ionic compounds
2. Enthalpy of atomization (Li): atomization of lithium to form one mole of gaseous ions
3. Bond enthalpy (F_2): Breaking one mole of F-F bonds to form one mole of fluorine atoms
4. Ionization energy: Ionization of lithium atom by removing one electron
5. Electron affinity: Adding one electron to fluorine atom to form fluoride ion
6. Lattice enthalpy: Breaking of one mole of solid lithium fluoride to form two gaseous ions

Experimental lattice enthalpy is determined through:

$$\Delta H^{\circ}_{lattice} = (-\Delta H^{\circ}_{f}) + \Delta H^{\circ}_{at}(alkali) + \Delta H^{\circ}_{at}(halogen) + \Delta H^{\circ}_{IE}(alkali) + \Delta H^{\circ}_{EA}(halogen)$$

Electronegativity difference

"Electronegativity is a measure of the ability of an atom to attract electrons in a covalent bond, and is described using the Pauling scale of values" [7] ΔElectronegativity is used to classify bonding characters, such as metallic, ionic, or molecular. If Δ> 1.8, the bond will have an ionic character. If Δ is between 0.5 and 1.8, the bond will have a polar covalent character. If Δ < 0.5, the bond will have a nonpolar covalent character[8].

Research questions

1. Is there a relationship between electronegativity difference and the difference between theoretical and experimental lattice enthalpy in monovalent ionic compounds using Kapustinskii and Born-Lande equation for theoretical value and Born Haber cycle for experimental value?
2. How does change in electronegativity difference of monovalent ionic compounds affect the difference between theoretical and experimental lattice enthalpies regarding the following factors in terms of impact in electrostatic attraction?
    1) Sum of ionic radii
    2) Bond length of ionic bond
    3) Bond strength of ionic bond

---

[6] Arrows for bond breaking points up, and arrows for bond forming points down.

[7] (Brown, 2014).

[8] (Brown, 2014).

Variables

Independent variables	- ΔElectronegativity - Sum of ionic radii of alkali metal and halogen - Bond length and strength of ionic bond
Dependent variable	- Absolute difference between theoretical and experimental lattice enthalpy values - Theoretical and experimental lattice enthalpy of monovalent ionic compounds
Controlled variables	- Monovalent ionic compounds formed with group 1 metals and group 7 halogens (MX lattice) - Experimental values obtained at standard condition (298.15K and 100kPa)

Methodology

Databases	a) IB chemistry data booklet (2016, Fourth edition) b) CRC Handbook of Chemistry and Physics c) Inorganic Chemistry: Principles of Structure and Reactivity
Resources	Microsoft excel

a) IB Chemistry data booklet: includes experimental lattice enthalpy values. As it is used in the IB curriculum, these data are highly reliable

b) CRC Handbook of chemistry and Physics: contains physical data for each ion and it is commonly used for the source of many other databases. It has been available for many years, and is cross-checked

c) Inorganic Chemistry: textbook published by Pearson Education that includes explanations and relevant data on lattice enthalpy. Since this source is a tertiary source, there is possible inaccuracy of data.

Data colleciton and processing

1. Collect raw data from the chosen databases.
2. Calculate Δ electronegativity between alkali metal and halogen ions, using obtained data from IB Data Booklet (2016).
3. Calculate the sum of ionic radii by adding their radius of ionic compound, using obtained data from IB Data Booklet (2016).
4. Substitute values into Kapustinskii Equation to obtain theoretical lattice enthalpy values.
5. Calculate average value for theoretical and experimental lattice enthalpies.
6. Calculate the absolute and %Δ between theoretical and experimental lattice enthalpies, using the average values obtained from step 5.
7. Create graphs with obtained data[9].
   i. ΔElectronegativity against absolute difference between theoretical and experimental lattice enthalpies sum of radii against theoretical and experimental lattice enthalpy values
   ii. Bond length against theoretical and experimental lattice enthalpy values
   iii. Bond strength against theoretical and experimental lattice enthalpy values[10]

---

[9] Graphing electronegativity difference against the percentage difference showed a relatively weaker relationship than that against the absolute difference. Therefore, I have decided to use the absolute difference values to find the correlation.

[10] This investigation is based on databases, so it is assumed that data is obtained in consideration of safety and ethical issues. Data collected electronically led to paper usage to be minimized., as well as saving the trees and the environment.

Results

Raw data

Table 1. Electronegativity and ionic radii of monovalent elements

Alkali metal (Group 1)			Halogens (Group 7)		
M.	Electronegativity	Ionic radii $(10^{-12}$ m$)$	X·	Electronegativity	Ionic radii $(10^{-12}$ m$)$
u·	1.0	76	F·	4.0	133
Na•	0.9	102	Cl-	3.2	181
K•	0.8	138	Br·	3.0	196
Rb•	0.8	152	1·	2.7	220
Cs+	0.8	167			

Going down groups 1 and 7, ionic radii increase due to increased energy levels. Moving down groups 1 and 7, the distance between the nucleus and valence electrons increases, resulting in decreased attraction, therefore lower electronegativity value. These values are the independent variables of the ionic compounds.

Processed data[11]

IF	ΔE	Σr	BL	BS	IF	ΔE	Σr	BL	BS
LiF	3.0	209	1.5639	577	KBr	2.2	334	2.8208	380
LiCl	2.2	257	2.0207	469	KI	1.9	358	3.0478	325
LiBr	2.0	272	2.1704	418	RbF	3.2	285	2.2703	494
LiI	1.7	296	2.3919	345	RbCl	2.4	333	2.7869	428
NaF	3.1	235	1.9260	519	RbBr	2.2	348	2.9447	381
NaCl	2.3	283	2.3609	412	RbI	1.9	372	3.1768	319
NaBr	2.1	298	2.5020	367	CsF	3.2	300	2.3454	519
NaI	1.8	322	2.7115	304	CsCl	2.4	348	2.9063	448
KF	3.2	271	2.1716	498	CsBr	2.2	363	3.0723	389
KCl	2.4	319	2.6667	433	CsI	1.9	387	3.3152	337

---

[11] Refer to footnotes for the abbreviations used in the first row of the table.

Table 3. Lattice enthalpies[12]

IF	$\Delta H_{lattice}$ **at 298K (kJmol^{-1})**										**ΔA**	**Δ%**
	TV					**EV**						
	K	**B**	**AV**	**SD**	**%U**	**IB**	**H**	**AV**	**SD**	**%U**		
LiF	960.4	1008.0	984.2	33.7	4.73	1049.0	1034.0	1041.5	10.6	1.43	57.3	5.50
LiCl	809.8	811.3	810.6	1.0	0.18	864.0	840.1	852.1	16.9	2.77	41.5	4.87
LiBr	772.7	766.1	769.4	4.7	0.86	820.0	781.2	800.6	27.4	4.73	31.2	3.90
LiI	717.5	708.4	713.0	6.4	1.27	764.0	718.4	741.2	32.2	5.97	28.3	3.81
NaF	872.8	902.0	887.4	20.6	3.24	930.0	914.2	922.1	11.2	1.70	34.7	3.76
NaCl	745.9	755.2	750.6	6.6	1.23	790.0	770.3	780.2	13.9	2.49	29.6	3.79
NaBr	713.3	718.8	716.1	3.9	0.76	754.0	728.4	741.2	18.1	3.40	25.1	3.39
NaI	666.6	661.2	663.9	3.8	0.81	705.0	680.7	692.9	17.2	3.45	29.0	4.18
KF	774.2	797.5	785.8	16.5	2.93	829.0	812.1	820.6	12.0	2.04	34.7	4.23
KCl	672.1	676.5	674.3	3.1	0.65	720.0	701.2	710.6	13.3	2.61	36.3	5.11
KBr	645.4	659.5	652.5	10.0	2.14	691.0	671.1	681.1	14.1	2.88	28.6	4.20
KI	606.8	623.0	614.9	11.5	2.60	650.0	632.2	641.1	12.6	2.74	26.2	4.09
RbF	741.4	761.1	751.3	13.9	2.59	795.0	780.3	787.7	10.4	1.85	36.4	4.62
RbCl	647.1	661.5	654.3	10.2	2.17	695.0	682.4	688.7	8.9	1.81	34.4	4.99
RbBr	622.3	636.4	629.4	10.0	2.21	668.0	654.0	661.0	9.9	2.10	31.6	4.79
RbI	586.3	602.5	594.4	11.5	2.69	632.0	616.7	624.4	10.8	2.42	30.0	4.80
CsF	709.2	723.0	716.1	9.8	1.91	759.0	743.9	751.5	10.7	1.99	35.4	4.71
CsCl	622.3	622.6	622.5	0.2	0.04	670	629.7	649.9	28.5	6.01	27.4	4.21
CsBr	599.3	599.6	599.5	0.2	0.05	647	612.5	629.8	24.4	5.33	30.3	4.81
CsI	565.8	568.2	567.0	1.7	0.42	613	584.5	598.8	20.2	4.65	31.7	5.30

---

[12] IF = Ionic formula
ΔE = Electronegativity difference (±0.2)
r = Sum of radii (±2pm)
BL = Bond length (±0.0001A)
BS = Bond strength (±1kJmol')
TV= Theoretical value (±0.1kJmol')
EV= Experimental value (±0.1kJmol')
K = Kapustinskii equation
B=Born-Land~ equation
AV = Average (±0.1kJmol ')
SD =Standard deviation (±0.1)
%U =Percentage uncertainty (±0.1)
1B =1B Data Booklet
H = Huheey
AA = Absolute difference (±0.1)
A% =Percentage difference (±0.1)

Comment on uncertainties

Database's calculation methods are not released, and although there must be some errors in each, the resources are reliable. Hence, uncertainties of each data correspond to the last decimal place of the value.

Sample calculations

Table 4. Sample calculations of Table 2 and 3

ΔElectronegativity = (electronegativity of halogen) - (electronegativity of alkali metal) E.g. LiF 4.0 - 1.0 = 3.0	Sum of ionic radii = (ionic radii of alkali metal)+ (ionicradii of halogen) E.g. LiF--+ 76pm + 133pm = 209pm	Average: $= \frac{\Sigma}{n}$ E.g. AV TV of LiF $\rightarrow \frac{960.4 + 1008.0}{2} = 984.2$
Percentage difference: $\frac{(AV\ EV)-(AV\ TV)}{AV\ EV} \times 100\%$ E.g LiF $\rightarrow \frac{1041.5-984.2}{1041.5} \times 100\% = 5.50\%$	Percentage uncertainty $= \frac{max\ value - min\ value}{max\ value} \times 100$ E.g. U for TV for LiF $\rightarrow \frac{1008.0 - 960.4}{1008.0} \times 100 = 4.275\%$	Range of %U = (m ax %U) - (min %U) TV 4.73% - 0.04% = 4.68% EV 6.01% - 1.43% = 4.58%
$\Delta H^{0}{}_{\text{lattice}} = K \cdot \frac{v \cdot \lvert z^{+}\rvert \cdot \lvert z^{-}\rvert}{r^{+} + r^{-}} \cdot \left(1 - \frac{d}{r^{+} + r^{-}}\right) \text{-} \times 10^{-3}$ E.g LiF $\rightarrow 1.20200 \times 10^{-4}\ J \cdot m \cdot mol^{-1} \cdot \frac{2 \cdot \lvert 1\rvert \cdot \lvert -1\rvert}{209\times10^{-12}m} \cdot \left(1 - \frac{3.45\times10^{-11}m}{209\times10^{-12}m}\right) \times 10^{-3} = 960.4$		Absolute uncertainty = I(AV TV)-(AV EV)I E.g. AU of LiF | 984.2 - 1041.5|= 57.3

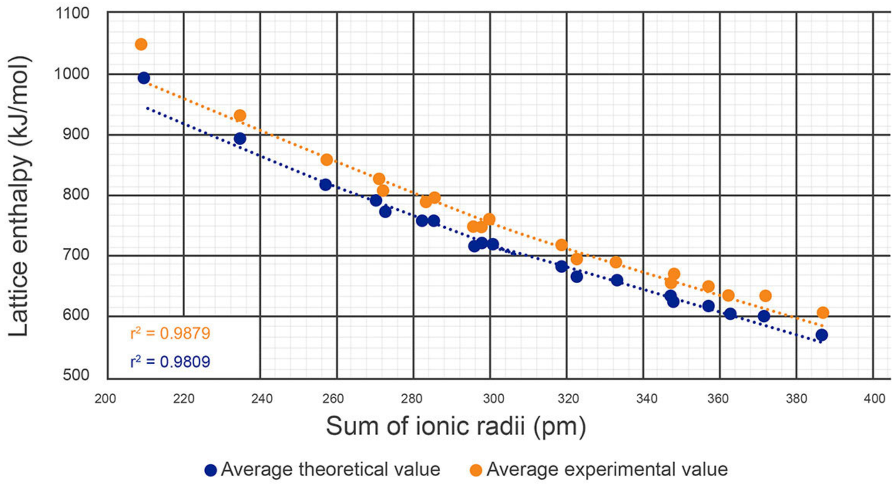
Evaluation of Reliability of Data

The data sources were evaluated by comparing the collected data. Table 3, shows a good agreement between the two average values. I calculated the absolute difference as uncertainty value. 4.68% and 4.58% of the percentage uncertainty range indicate the values are close to the average values. Thus, results are reliable and can be used to show a correlation in lattice enthalpies of each ionic compound.

Analysis and conclusion

1) Effect of sum of ionic radii

Figure 2. Sum of ionic radii



[13]

---

[13] Error bars are not included in Figure 2 because it would be skewed by the scattered values.

Figure 2, shows that as the sum of ionic radii of metal and halogen ions increases, the lattice enthalpy decreases. As the distance increases, the attraction between them will decrease, as the nuclei are further apart, and less energy is required to break one mole of solid ionic compound to form gaseous ions (decrease in lattice enthalpies). The negative relationship is not linear, and appears to be steeper at lower sum of radii and the same increase in distance between two nuclei has less effect as the distance increases.

Initially, the decrease is more significant than the same decrease in the ionic radii when the radius is doubled. The gap between theoretical and experimental values seems to narrow down as sum of ionic radii increases, suggesting that very large ionic radii would have smaller differences. The differences would be minimized only if the sum of radii value is very high and there would be still considerable differences between the TV and EV even if the sum of ionic radii is 0.

Uncertainties of ionic radii are within 5 % as radii values are too small. Data are significant and errors are too small to be observed. Assuming that error bar values are uncertainty values of each TV and EV, the biggest error bar would be with LiF (46.6) for TV, and CsCl (39.1) for EV. Overall, the sum of ionic radii of monovalent ionic compounds as a factor of determinant of electronegativity may contribute to change in the gap between theoretical and experimental lattice enthalpies.

2) Effect of bond length of ionic bond

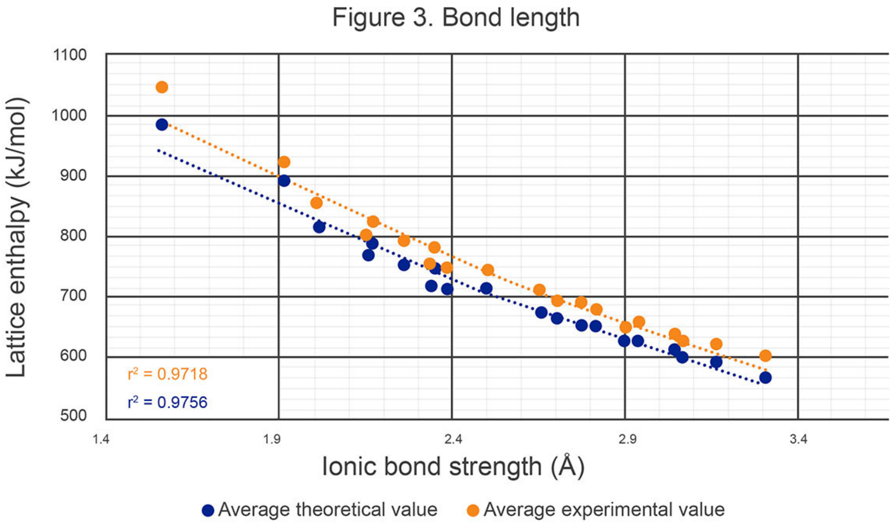


Figure 3 shows that the average theoretical lattice enthalpy values of ionic compounds are constantly higher than the corresponding average experimental lattice enthalpy values. The gradients of two trend lines are similar and the distance of theexponential equations are only 48.4 apart. There is a pattern in ionic bond length and lattice enthalpies as the theoretical and experimental values tend to decrease at the same rate. This may result from the relationship between ionic bond length and the electrostatic forces of ionic compounds, suggesting that ionic bond length has a negative correlation with lattice enthalpy values, increasing ΔElectronegativity of ionic compounds.

I measured the sum of ionic radii and bond length despite those having approximately the same values, because the *"ionic character of a bond depends on the strength of the force holding ions in lattice. Hence, the stronger the force of attraction, the greater the lattice enthalpy"*[14] Ionic radii only take account of ionic character, so I compared ionic bond lengths to attempt identifying new relationships in theoretical lattice enthalpies which are considered to exclude any ionic characters.

Uncertainties are within 5% indicating that values are significant with high reliability. Overall, decreasing trend in ionic bond length corresponding to lattice enthalpies suggests that electronegativity values may cause change in the difference between theoretical and experimental lattice enthalpies.

---

[14] (Brown, 2014).

3)Effect of bond strength of ionic bond

Figure 4. Bond Strength

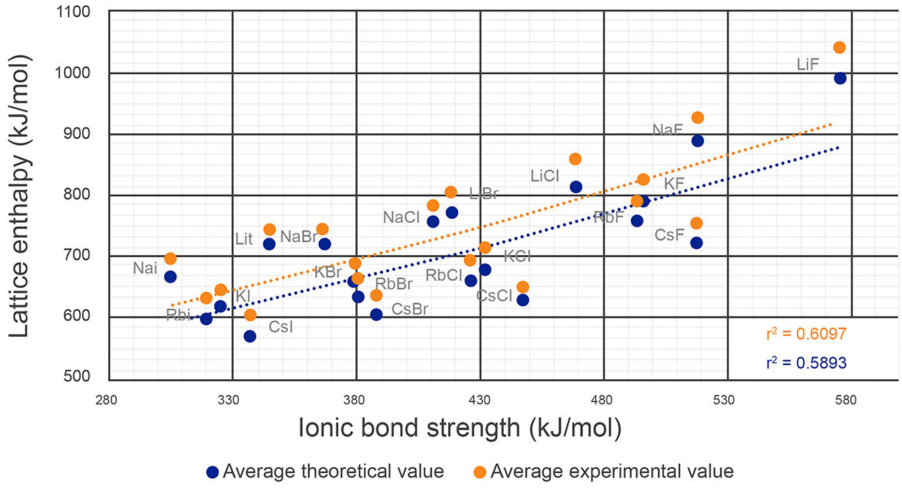
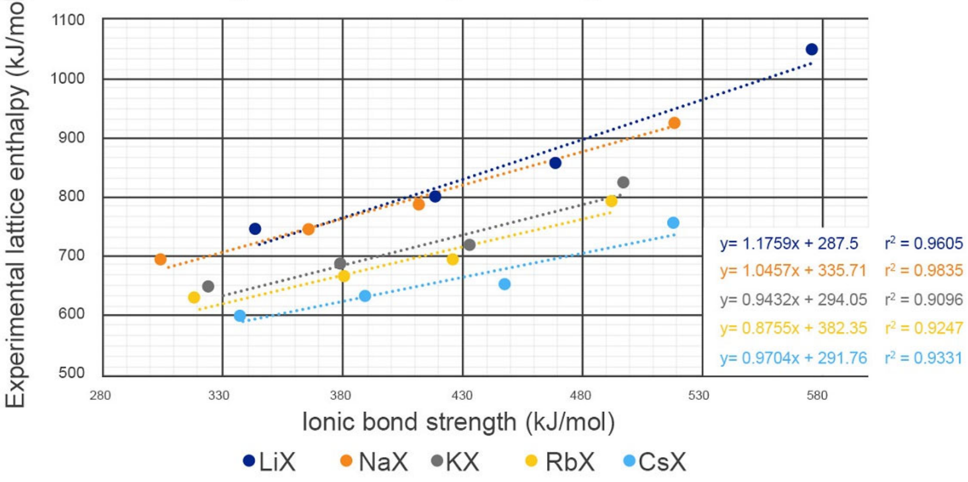


Figure 4 shows a trend but it is not very well defined. r^2 values of theoretical and experimental values (0.6097 and 0.5893, respectively) are too low to conclude a strong relationship, but suggest that for an increase in bond strength, there is also an increase in the other. Both of the exponential expressions exhibit similar gradients and y-intercepts, suggesting that there will be slightly decreasing distance between those exponential trendlines as they converge with increasing bond strength. Error bars are added, but not possible to be observed.This represents that data are reliable with small variance.

Although the trend-lines of the different ionic compounds do not converge as clearly as with the factors above, there is relatively weak correlation between bond strength and average theoretical and experimental lattice enthalpies, but no definite or strong correlation between bond strength and electronegativity value. Polarity tends to increase bond strength as the opposite charges attract each other, forming stronger bonds due to higher electronegativity difference (Libretexts, 2020). Additional strength from covalent character of ions is found in experimental values as theoretical only take account of the ionic model (Jones, 2014). It is hard to state that bond strength can directly affect the electronegativity value that can change the gap between theoretical and experimental lattice enthalpies, even though both of the lattice enthalpy values have positive exponential trendlines that are farther apart with increasing bond strength. I may infer that bond strength has insignificant impact on electronegativity values, leading to weak correlations in the difference between theoretical and experimental lattice enthalpies.

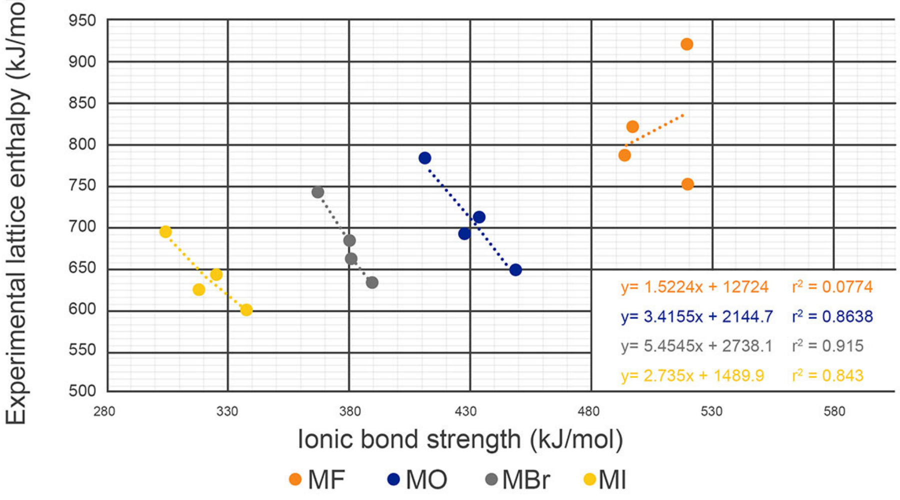
On further inspection, I wondered if there are patterns shown in the periodic table, similar to the trend of ionization energy down group 7 and group 1. I labeled the data on Figure 4 and found an increasing trend in lattice enthalpy values moving down group 7. There was general decreasing order moving down group 1 and there was a slight increase in lattice enthalpies from Rb to K, while Li showed different patterns.

Figure 5. Bond Strength according to metals



I plotted another diagram using experimental lattice enthalpy values as dependent variables to take ionic characters into account. There were positive trends as moving down the group 7, and regarding group 1 , there is a slight decrease in the gradients of linear regression lines. This suggests that although there is the same increase in bond strength of Li, there will be a slight difference in that of CsX. This may be due to the size of difference in between group 1 which can affect the bond strength and lattice enthalpies.

Figure 6. Bond Strength according to halogens



I plotted values moving down the group such as MF, MCl, MBr, and MI[4]. There were generally decreasing trends, except for MF series. This may be due to the small size of Fluoride affecting both bond strengths and experimental lattice enthalpies.

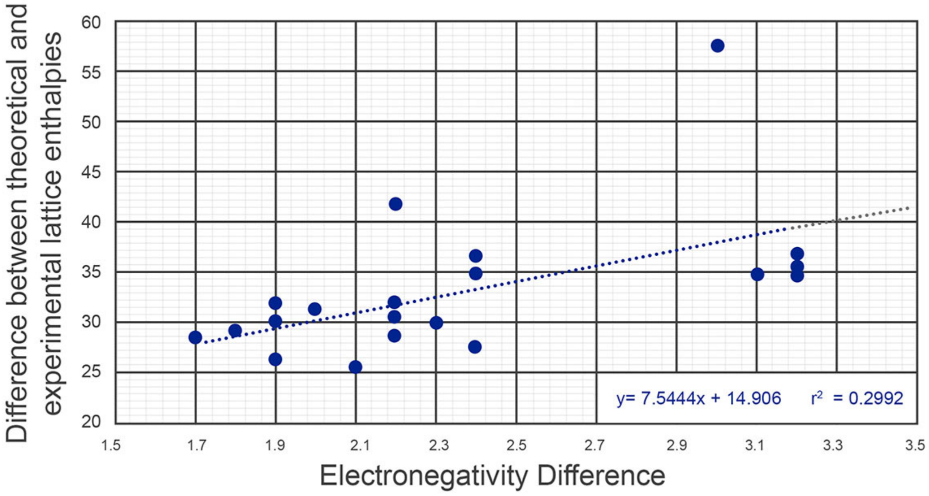
There was no clear pattern in the gradients, especially for MF compounds, as values don't decrease as you move down the group 7. MF compounds exhibit higher lattice enthalpy values than other compounds. We can state that increasing size of the halogen corresponds with lower lattice enthalpy. Moving down group 1, lattice enthalpy decreases. Despite not consistent trends in metal halogen ionic compounds, I may state that there is a general negative trend observed in increasing bond strength and experimental lattice enthalpy moving down group 1 metals. The ionic compounds tend to exhibit general trends moving down group 1 metals and group 7 halogens, although there is not a clear correlation between bond strength and electronegativity.

3. Effect of ΔElectronegativity[15]

---

[15] I have excluded Li compounds' values as they tend to not follow the pattern of the other group 1 metals. I believe this is due to its smallest size that results in the highest bond strength and lattice enthalpy values.

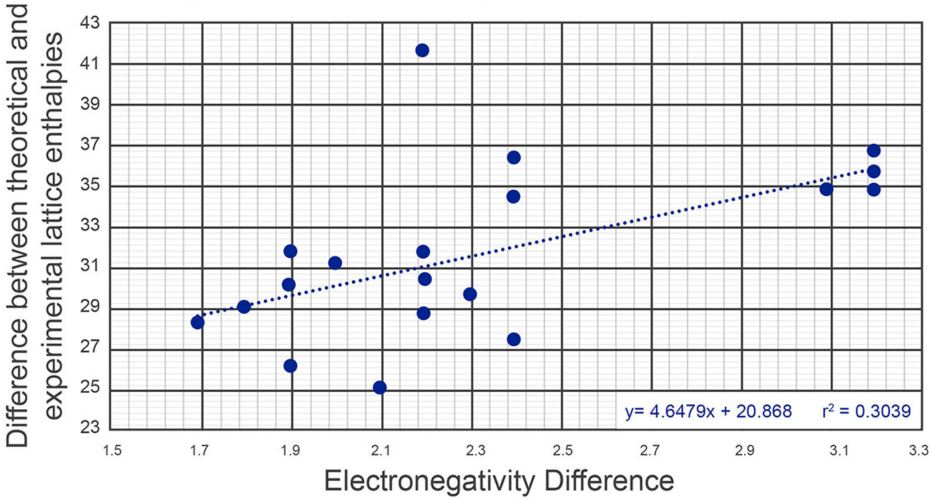
Figure 7. Electronegativity difference



This graph summarizes the effect of factors affecting electronegativity values and answers my hypothesis. Due to minimal uncertainties of the absolute difference between theoretical and experimental lattice enthalpy (±0.1), y-value error bars are not observed.

The result goes against my hypothesis. There is a weak correlation and we can observe few data not following the trend, especially LiF compounds. This indicates that there is not really a close relationship between those variables, probably due to procedural errors in enthalpy measurement, such as heat loss from the system or possible pressures of the gases involved (Jason, 2017).

Figure 8. Electronegativity difference (Excluding LiF)



As in Figure 7, there is a slight increase in r^2 value (0.3039), but the value is too low, and there is not much relationship between ΔE and ΔA. Overall, I can conclude that there is no relationship between electronegativity difference and the difference between theoretical and experimental lattice enthalpies.

In brief:

1) The sum of ionic radii, bond length, and bond strength of ionic compounds formed from alkali metals and halogens show higher lattice enthalpies on theoretical values than experimental values. This is because those affect electronegativity values in ionic compounds that can affect the difference between theoretical and experimental lattice enthalpies. The relationships in sum of ionic radii and bond length were relatively stronger than that of bond strength. Instead, there were trends related to periodic characters observed in the relationship with bond strength.
2) The effect of electronegativity differences of ionic compounds on changes of difference in theoretical and experimental lattice enthalpies is inconsiderable. As the electronegativity increases the change in difference between theoretical and experimental lattice enthalpy is less responsive in terms of covalent character of ionic compounds.

My hypothesis was partially refuted. Extra conditions should be also considered in order to obtain a clear correlation between ΔElectronegativity and the difference between theoretical and experimental lattice enthalpies.

Evaluation

Limitation	Significance	Improvement
Different methods of measuring experimental data	The difficulties in controlling heat loss and standard conditions might have contributed to further inaccuracy of data (systematic error).	It would be better to use data from one specific database that has a uniformed method of obtaining experimental data. This would reduce the variance in experimental data and increase the reliability of obtained data.
Uncertainty values from each database	As uncertainty values are not mentioned from the databases that I used, it is difficult to calculate exact uncertainty values, lowering the reliability of data.	Finding the primary source of data is essential to calculate the exact uncertainty values from each database. As major uncertainty values come from volumetric apparatus, the exact methodology for each database can be found and applied in uncertainty calculation.
Exclusion of ionic compounds that are not formed with alkali metals and halogens	I have excluded ionic compounds that are not formed with group 1 and group 7 from my investigation although these are monovalent ionic compounds due to unobtainable data. For example, adding AgX compounds to investigation would have showndifferences in electronegativity, leading to different conclusions that what I have got.	Comparing the values and trends of other monovalent ionic compounds that are not formed with alkali metals and halogens with the values I found in this investigation would lead to more thorough and depth study. This can make me understand the reason behind some of the extreme values and the weak relationship between electronegativity differenceand the difference between theoretical and experimental lattice enthalpies.
Subjective selection of collected data	As I mainly collected my data from 1B Chemistry data booklet and CRC Handbook of Chemistry and Physics, there might be possible inaccuracy on percentage difference derived from each source used to obtain lattice enthalpies.	I could use average value from data collected from five different databases for each variable would be similar with having five trials in experimental investigation. This would reduce the error and increase reliability.
Lack in representation of reality	Experimental lattice enthalpies are obtained via Born-Haber cycles and require long series of steps.Theoretical lattice enthalpies might be approximate value instead of the exact value, which may lack representation of reality.	It is only possible toreduce the errors found in experimental lattice enthalpies, but not realistic at school level.

References

“Born–Landé Equation.” *Wikipedia*, Wikimedia Foundation, 24 May 2020, en.wikipedia.org/wiki/Born%E2%80%93Land%C3%A9_equation.

Brown, Catrin, and Mike Ford. *Higher Level Chemistry: for The IB Diploma*. Pearson, 2014.

“Chemistry Data Booklet- IB Chem Notes for New 2016 Syllabus.” *International Baccalaureate Organization 2014*, 2016, ibalchemy.com/data_booklet_2016.pdf.

Huheey, James E., et al. *Inorganic Chemistry: Principles of Structure and Reactivity*. HarperCollins, 1993.

Jason. “Effects of Experimental Errors on Enthalpy Measurements …” *Youtube*, 2017, www.youtube.com/watch?v=MyFfc-vIDNs.

Jones, Mike Sugiyama. “15.2.4 Discuss Differences between Theoretical …” *Youtube*, 2014, www.youtube.com/watch?v=nd7jIP9yaLE.

“Kapustinskii Equation.” *Wikipedia*, Wikimedia Foundation, 2 Oct. 2020, en.wikipedia.org/wiki/Kapustinskii_equation.

Libretexts. “8.4: Bond Polarity and Electronegativity.” *Chemistry Libre Texts*. Libretexts, 14 July 2020, chem.libretexts.org/Bookshelves/General_Chemistry/Map:_Chemistry_-_The_Central_Science_(Brown_et_al.)/08_Basic_Concepts_of_Chemical_Bonding/8.4:_Bond_Polarity_and_Electronegativity.

Lide, David R., editor. *CRC Handbook of Chemistry and Physics*. Taylor & Francis, 2003.

Shannon, Robert D. “Revised Effective Ionic Radii and Systematic Study …” *ResearchGate*, 1976, https://www.researchgate.net/publication/220037484_Revised_Effective_Ionic_Radii_and_Systematic_Study_of_Inter_Atomic_Distances_in_Halides_and_Chalcogenides.