<table>
<thead>
<tr>
<th>Lesson</th>
<th>Topics Covered</th>
<th>Reference</th>
<th>Homework/Assignments</th>
</tr>
</thead>
</table>
| 5 & 6  | Take up Questions from Review  
Introduction to Bonding  
Type of Bonding depends on IE and EN  
• Ionic Bonding:  
  metals + non-metals (ΔEN >1.70)  
  review bonding and properties from reading  
• Covalent Bonding:  
  non-metals + non-metals (ΔEN 0.00 to 1.70)  
  see later lessons  
• Metallic Bonding  
  metals + metals  
  the “Free Electron Model” of metallic bonding  
Properties of Metallic Substances  
Metal Crystal Structures  
Alloys  
• substitutional and interstitial | Note: Introduction to Bonding  
Handout: Review of Ionic Bonding  
Text: p. 165 – 166  
Handout: Metal Crystal Structures  
Text: p. 170 - 171 | Study for Quiz  
Homework on Ionic Compounds:  
1. Read handout: Review of Ionic Bonding  
2. Page 165, Q 1 – 3 (just do electron dot diagrams for Q3)  
3. Write ionization reactions (without electron configurations) to show the formation of the ions: Mg$^{2+}$, O$^{2-}$, Sc$^{3+}$, Si$^{4+}$, Si$^{4-}$  
3. Predict 4 properties of RbBr$_2$. |
| 6      | Quiz on Chapter 3 (30 minutes)  
Finish Metallic Bonding | Handout: Metal Crystal Structures  
Text: p. 170 - 171 | Homework questions on handout: Metal Crystal Structures |
| 7 & 8  | Introduction to Covalent Bonding  
• non-metals + non-metals  
Drawing Covalent Compounds  
Hybridization Theory (sp$^3$)  
Polarity in Bonding  
The Bonding Continuum  
% Ionic Character  
Properties of Covalent Compounds | Note: Introduction to Covalent Bonding and the Octet Rule  
Text: p. 167-169 | Homework questions at the end of the note: Introduction to Covalent Bonding and the Octet Rule |
| 9      | Variations of Covalent Bonding by the Octet Rule  
• the “System” for bonding by the Octet Rule  
• Co-ordinate covalent bonds  
• Resonance structures | Note: Variations on Covalent Bonding by the Octet Rule  
Handout: The “System” for Bonding by the Octet Rule  
Text: Page 173-177 | Homework questions at the end of the handout: The “System” for Bonding by the Octet Rule |
| 10     | Variations of Covalent Bonding, Exceptions to the Octet Rule  
• hybridization theory (revisited: sp and sp$^2$)  
• insufficient electrons to complete a stable octet  
• expanded valence (electron promotion) | Note: Variations of Covalent Bonding, Exceptions to the Octet Rule  
Handout: The “System” for Expanded Valence  
Text: Page 177-178 | Homework questions at the end of the handout: The “System” for Expanded Valence |
### Unit #1, Chapter 4 Outline (Lessons 10 – 14)
#### Structure and Properties of Substances

| 11 & 12 | VSEPR Theory: Shapes of Molecules | Note: VSEPR Theory | 1. Homework at the end of VSEPR Summary handout  
2. Begin VSEPR Assignment: Predicting the Shape and Polarity of Molecules.  
Due: ____________ |
| --- | --- | --- | --- |
| a) less than a stable octet:  
  i) \(AX_2E_0, AX_3E_0, AX_2E_1\)  
 b) stable octet:  
  i) \(AX_4E_0, AX_3E_1, AX_2E_2\)  
 c) expanded valence:  
  i) \(AX_5E_0, AX_4E_1, AX_3E_2\)  
  ii) \(AX_6E_0, AX_5E_1, AX_4E_2\) | Handouts:  
1. VSEPR Theory and Molecular Shape  
2. VSEPR Summary  
3. VSEPR Assignment | Text: Page 178 - 186 |

| 13 | Determining the Polarity of Molecules | Note: Polarity of Molecules | 1. Page 188: Q 23 – 26  
3. Work on VSEPR Assignment |
| --- | --- | --- | --- |
| • symmetry of molecules  
• polarity of bonds | Handout: Intra- and Inter-molecular forces of attraction | Text: Page 187 – 189 |

| 14 | Intra-molecular Forces  
Inter-molecular Forces of Attraction | Handout: Intra- and Inter-molecular forces of attraction | 1. Study chart on page 195.  
2. Page 208: Q 1, 2, 6, 7, 8  
3. Complete VSEPR Assignment: Predicting the Shape and Polarity of Molecules. |
| --- | --- | --- | --- |
| • chemical bonds (ionic and covalent)  
**Inter-molecular Forces of Attraction**  
• Ion-ion attraction  
• Dipole-dipole forces (dip-dips)  
• Hydrogen bonding (H-bonding)  
• London dispersion forces (LDFs) | Complete Handout from Lesson 11: Determining the Polarity of Covalent Molecules | Text: Page 190 – 195  
Omit sections on ion-dipole and induced inter-molecular forces on page 191. |

| 15 | Unit Summary: Classifying Solids  
**Crystalline Solids**  
• Ionic solids  
• Network solids (covalent)  
  - allotropes  
• Metallic solids  
• Molecular solids (covalent)  
  a) polar covalent  
  b) pure covalent  
**Amorphous Solids (non-crystalline solids)**  
Prelab for Lab #1: Using Physical Properties to Classify Solids | Handout: Summary: Characteristics of Crystalline Solids  
2. Complete Handout: Summary of Characteristics of Crystalline Solids  
Read through Lab handout in preparation for Lab tomorrow. Come to class with your observation chart made.  
Begin to study for unit test (Chapters 3 and 4). Review for Chapter 4 on Internet.  
**Unit 1 Test:** ____________ |
| --- | --- | --- | --- |

<table>
<thead>
<tr>
<th>16</th>
<th>Lab #1: Using Physical Properties to Classify Solids</th>
<th>Text: Page 196 - 204</th>
</tr>
</thead>
</table>
| 1. Begin Lab write up.  
Due: ____________  
2. Study for unit test (Chapters 3 and 4). Review for Chapter 4 on Internet. |

| 17 | Take up questions from Lab, Review  
Begin Organic Chemistry  
1. Intro and definitions  
2. Alkanes |  |  |
| --- | --- | --- | --- |
Ionic bonding usually occurs between metal and non-metal elements, as follows:

- metal atoms have low electronegativity and low ionization energies
- they tend to lose valence electrons to achieve a more stable electron arrangement
- by losing electrons, many metals form positive ions (cations) which are isoelectronic with a Noble gas

\[
\begin{align*}
\text{eg. } & \text{ Na} \rightarrow \text{Na}^{1+} + \text{e}^- \quad \text{Because metals lose electrons, show the electron on the product side (right) of the ionization equation} \\
\text{eg. } & \text{Ca} \rightarrow \text{Ca}^{2+} + 2 \text{e}^- \\
\text{eg. } & \text{Al} \rightarrow \text{Al}^{3+} + 3 \text{e}^-
\end{align*}
\]

- non-metals have high electronegativity and high ionization energy
- they tend to gain valence electrons to achieve a more stable electron arrangement
- by gaining electrons, non-metals form negative ions (anions) which are isoelectronic with a Noble gas

\[
\begin{align*}
\text{eg. } & \text{Cl} + \text{e}^- \rightarrow \text{Cl}^{1-} \quad \text{Because non-metals gain electrons, show the electron on the reactant side (left) of the ionization equation} \\
\text{eg. } & \text{S} + 2\text{e}^- \rightarrow \text{S}^{2-} \\
\text{eg. } & \text{N} + 3\text{e}^- \rightarrow \text{N}^{3-}
\end{align*}
\]

- positively charged metal ions are attracted to negatively charged non-metal ions by electrostatic attraction (the attraction between oppositely charged objects)
- the electrostatic attraction between cations and anions forms an ionic bond
- ionic bonding involves a substantial (or nearly complete) transfer of electrons from one atom to another
- ionic bonding usually occurs between a metal and a non-metal, when the difference in electronegativity between the bonding atoms is greater than 1.70

We can show the formation of ionic compounds using electron dot (Lewis dot) diagrams.
Follow these steps:
1. draw the electron dot diagrams for each element as neutral atoms
2. draw arrows to show the direction that the electrons will move
3. draw the ions with their charges inside square brackets in the ratio that is required so that the number of electrons being lost by the metal atom(s) is equal to the number of electrons being gained by the non-metal atom(s). Show the ions in an arrangement that spreads out the positive and negative ions.
4. write the chemical formula for the ionic compound that will form

eg. Show the formation of the ionic compound that forms between magnesium and oxygen

\[
\text{Mg} \rightarrow \text{O} \quad \Rightarrow \quad [\text{Mg}]^{2+} + [\text{O}^2-] \quad \Rightarrow \quad \text{MgO}
\]

eg. Show the formation of the ionic compound that forms between sodium and sulfur

\[
\text{Na} \rightarrow \text{S} \quad \Rightarrow \quad [\text{Na}]^{1+} + [\text{S}^2-] + [\text{Na}]^{1+} \quad \Rightarrow \quad \text{Na}_2\text{S}
\]
The elements in ionic compounds combine in ratios such that the number of electrons that are lost by the metal must equal the number of electrons that are gained by the non-metal. Similarly, in an ionic compound, the total number of positive charges equals the total number of negative charges. That is, while the compound is made up of positive and negative ions, the overall charge of the compound is neutral.

**Properties of Ionic Compounds**

When metal and non-metal elements react to form positive and negative ions, every negative ion is attracted to every positive ion, and visa versa. Because of these multiple attractions, the ions arrange themselves in a “crystal lattice” pattern of alternating positive and negative ions.

The ions arrange themselves as close to each other as possible, which gives them the lowest possible potential energy. The tremendous amount of energy that is **released** as the ions arrange themselves into a crystal lattice is called “lattice energy”.

In the crystal lattice, there is never a single “LiCl” molecule. All of the Li\(^{1+}\) ions are attracted to all of the Cl\(^{1-}\) ions. The formula unit “LiCl” tells us that ratio of lithium ions to chlorine ions in the crystal lattice is one to one.

The attraction between the formula units (‘molecules’) of ionic compounds is very strong. Remember that the attraction between molecules is called “**inter-molecular** attraction” (“inter” means between) while the attraction within molecules (the chemical bonds that hold the atoms together in a compound) is called “**intra-molecular attraction**”.

If we consider one formula unit of NaCl to be one ‘molecule’, then the ionic bond between the Na\(^{1+}\) and Cl\(^{1-}\) ions is the intra-molecular attraction. The attraction between adjacent NaCl ‘molecules’ is the inter-molecular attraction. The diagram below may help:

Because of the strong inter-molecular attraction in ionic compounds, they have the following physical properties:
1. all ionic compounds are solid at room conditions (SATP)
2. they have high melting and boiling points (it requires a lot of energy to disrupt the crystal lattice)
3. they are hard, crystalline solids
4. they usually have very little odour because it requires a huge amount of energy for formula units (‘molecules’) of ionic compounds to separate themselves from the crystal lattice to be able to travel through the air and land in receptors in our noses to create a smell.
Unit 1, Lesson 05: Metal Crystal Structures

1. Simple cubic packing - the layers of atoms are arranged directly over one another. This is the most uncommon and least dense arrangement of atoms. It leaves large spaces between the atoms.

2. Body-centred cubic packing - the second layer fits into the spaces over the first layer. The third layer fits into the spaces over the second layer and is directly above the first layer. This is the most dense arrangement of atoms.

3. Face-centred cubic packing - the second layer has four atoms while the layers above and below it have five atoms. This is the second most dense arrangement of atoms.

4. Hexagonal packing: the atoms form a three-dimensional hexagonal shape. The second layer has six atoms arranged in a hexagon around a single central atom, while the layers above and below it have three atoms that fit into the spaces above and below the central layer. This is also quite dense.

Metallic Bonding in Alloys

An alloy is a homogeneous mixture (a solid in solid solution) of a base metal with other elements (other metals or non-metals such as carbon or silicon). The base metal is the metal which is present in the largest concentration. The elements are mixed when they are molten (melted) and form metallic bonds as they solidify to form an alloy.

eg. sterling silver is an alloy of 92.5% silver and 7.5% copper, by weight.

eg. bronze is an alloy of (about) 75% copper and 25% tin.

Substitutional Alloys:
If the added metal atoms are about the same size as the base metal atoms, then the added atoms just mix evenly through the alloy and the crystal structure remains similar to that of the pure base metal. This is called a “substitutional alloy”.

eg. brass, made of copper and zinc

eg. stainless steel, made of iron, chromium and nickel
Interstial Alloys
If the added atoms are much smaller than the base metal atoms, they can fit into the spaces between the base metal atoms. This is called an “interstitial alloy”.

eg. high-carbon steel, made of iron (Z = 26) and carbon (Z = 6)

eg. an alloy containing tungsten (Z = 74) and cobalt (Z = 27)

Properties of Alloys
Both types of alloys are usually much harder and stronger than pure metals because of TWO factors:
1. the added metal may provide additional valence electrons which will increase the strength of the metallic bonding
2. the different sizes of the added atoms will interfere with how easily the metal atoms can slide past one another, making the metal stronger and less malleable

Metallic Bonding Homework:
1. Read pages 170 to 171.
2. Describe metallic bonding.
3. The electrons in a piece of metal are described as “delocalized”. Explain what this means.
4. Explain why metals are good conductors of electricity.
5. Explain why metals are malleable and ductile.
6. Explain why titanium (used in aircraft) is harder and stronger than calcium.
7. Use your periodic table to find two elements with each of the following crystal structures:
   a) (simple) cubic: ___________________ and ___________________
   b) body centred cubic: ___________________ and ___________________
   c) face centred cubic: ___________________ and ___________________
   d) hexagonal: ___________________ and ___________________

8. Define alloy.
9. Brass is an alloy of about 85% copper and 15% zinc. How would the properties of brass compare to the properties of pure copper? Explain.
10. Razor blades are often coated with an alloy of chromium and platinum to increase hardness and keep the blade sharp longer. Is this a substitution or interstitial alloy? Explain.
11. One type of steel used to make railway tracks is an alloy of iron, chromium, and vanadium. Is this a substitution or interstitial alloy? Why isn’t pure iron used for railway tracks?

To help you visualize the arrangement of atoms in different crystal structures:
You can get an idea of the crystal structures of metals and alloys by modeling with pennies.
What type of crystal structure does each of the following penny arrangements model?
Covalent compounds are made of uncharged or only partially charged atoms so there is relatively low inter-molecular attraction between the particles in covalent compounds.

Because of the weak inter-molecular attraction in covalent compounds, they have the following physical properties:
- may be gases, liquids or soft solids at room conditions (SATP)
- have low melting and boiling points
- often have distinctive odours (because molecules separate easily from one another to travel through the air to land in receptors in our noses and create a smell)
- are non-conductors of electricity in their pure form (the electrons are bonded between atoms and are not free to move from atom to atom)
- usually do not conduct electricity in solution (they are non-electrolytes)

**Homework:**
1. Read pages 167 to 169.
2. Define bond energy (page 168). What does bond energy measure?
3. Which has higher bond energy: C – C, C = C, or C≡C?
4. Do Questions 6, 7 and 8 on pages 169-170 of McGraw-Hill. For question 8, use our criteria for pure and polar covalent bonds ($\Delta EN < 0.50$ and $\Delta EN$ between 0.50 – 1.70, respectively)
5. Complete the chart below to classify each bond as pure covalent, polar covalent or ionic. Use our criteria for pure and polar covalent bonds ($\Delta EN < 0.50$ and $\Delta EN$ between 0.50 – 1.70, respectively):

<table>
<thead>
<tr>
<th>Bonded Atoms</th>
<th>$\Delta EN$</th>
<th>% Ionic Character</th>
<th>Type of Bond</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) H – Cl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>b) Cl – F</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>c) Cl – Cl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>d) Na – Cl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>e) C – S</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>f) Cs – I</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>g) Fr – F</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>h) N – I</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>i) F – O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>j) C – O</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

6. Use the Octet Rule to draw the structural diagrams (aka Lewis or ‘stick’ diagrams) for the following covalent molecules. Recall the general rules from Grade 11:
- oxygen does not usually bond to oxygen except in O$_2$, O$_3$ and H$_2$O$_2$ and
- if there are two or more atoms of carbon in a molecule, they are usually bonded to each other

<p>| | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>a) F$_2$</td>
<td>b) HI</td>
<td>c) N$_2$</td>
<td>d) SiS$_2$</td>
</tr>
<tr>
<td>f) CBr$_4$</td>
<td>g) CCℓ$_2$F</td>
<td>h) HCCℓ$_3$</td>
<td>i) C$_2$H$_6$</td>
</tr>
<tr>
<td>k) C$_2$H$_2$</td>
<td>l) CH$_2$NH</td>
<td>m) N$_2$H$_4$</td>
<td>n) N$_2$H$_2$</td>
</tr>
</tbody>
</table>

7. Explain what is meant by hybridized orbitals. Draw the predicted and actual orbital diagrams for silicon. Explain why the hybrid orbitals are named sp$^3$.

8. In chart form, compare the $\Delta EN$ and typical physical properties of ionic and covalent compounds. How are these properties related to $\Delta EN$?
These steps may only use for bonding by the octet rule! All atoms (except H) achieve a stable octet. This system MUST be followed when the central atom is C, N, or O.

<table>
<thead>
<tr>
<th>Steps</th>
<th>Example: BrO$_4^{1-}$ (perbromate ion)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. Determine the central atom:</td>
<td>Central atom:</td>
</tr>
<tr>
<td>• the first atom in the chemical formula (the lowest EN) goes in the middle, except hydrogen</td>
<td></td>
</tr>
<tr>
<td>2. Calculate the # of valence electrons we HAVE:</td>
<td>We have:</td>
</tr>
<tr>
<td>• total up the number of valence electrons available on the unbonded atoms</td>
<td></td>
</tr>
<tr>
<td>• add or subtract e- to account for the charge on ions</td>
<td></td>
</tr>
<tr>
<td>3. Calculate the # of electrons we NEED:</td>
<td>We need:</td>
</tr>
<tr>
<td>• use the octet rule to find the number of electrons needed to give all atoms a stable octet; hydrogen is stable with just 2 electrons</td>
<td></td>
</tr>
<tr>
<td>4. Calculate the # of electrons that are in bonds:</td>
<td>e- in bonds:</td>
</tr>
<tr>
<td>• subtract the number of electrons we HAVE (step 2) from the number if electrons we NEED (step 3)</td>
<td></td>
</tr>
<tr>
<td>5. Calculate the # of bonds that must form:</td>
<td># of bonds:</td>
</tr>
<tr>
<td>• divide the # of electrons in bonds (step 4) by 2 (each bond is made of 2 electrons)</td>
<td></td>
</tr>
<tr>
<td>6. Draw a skeleton structure:</td>
<td>Skeleton structure:</td>
</tr>
<tr>
<td>• draw the central atom and bonded atoms</td>
<td></td>
</tr>
<tr>
<td>• distribute the bonds (from step 5) to attach all atoms</td>
<td></td>
</tr>
<tr>
<td>7. Calculate the # of lone pairs (unshared e-):</td>
<td># of LP</td>
</tr>
<tr>
<td>• subtract the # of electrons that are in bonds (step 4) from the # of electrons that we have (step 2)</td>
<td></td>
</tr>
<tr>
<td>• divide the # of electrons by 2 to find the number of lone pairs (LP)</td>
<td></td>
</tr>
<tr>
<td>8. Complete the Lewis structure:</td>
<td>Lewis structure:</td>
</tr>
<tr>
<td>• add the electrons in lone pairs (step 7) around the atoms to complete stable octets for all atoms (except H)</td>
<td></td>
</tr>
<tr>
<td>• for ions, draw brackets around the finished structure and include the charge on the ion</td>
<td></td>
</tr>
</tbody>
</table>

**Homework:**
1. Read pages 173 to 177.
2. On page 177, questions 9 – 13. For questions 9 and 12, you do not need the “system”.
3. Draw Lewis structures for ozone (O$_3$), BrFO$_3$, NO$_3^{1-}$, NH$_4^{1+}$, NO$_2^{1-}$, CO, C(F$_2^{1+}$, SeO$_3^{2-}$. Include any resonance structures.
These steps should be used when the octet rule will not work because the number of bonds required to provide each atom with a stable octet is not sufficient to bond all of the atoms together into a molecule.

Only the central atom can have more than a stable octet!!!

<table>
<thead>
<tr>
<th>Steps</th>
<th>Example: CℓF₄⁺</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. <strong>Determine the central atom:</strong></td>
<td>Central atom:</td>
</tr>
<tr>
<td>• the atom with lowest EN (except H) goes in the middle</td>
<td></td>
</tr>
<tr>
<td>2. <strong>Determine if the central atom has an expanded valence.</strong></td>
<td>Is it octet rule or expanded valence?</td>
</tr>
<tr>
<td><strong>The Rules:</strong></td>
<td></td>
</tr>
<tr>
<td>• C, N, O and F can NOT have expanded valence</td>
<td></td>
</tr>
<tr>
<td>• if P has more than 3 bonds, it can be expanded</td>
<td></td>
</tr>
<tr>
<td>• if S has more than 2 bonds, it can be expanded</td>
<td></td>
</tr>
<tr>
<td>• if a Cl, Br or I has more than 1 bond, it can be expanded</td>
<td></td>
</tr>
<tr>
<td>• if a Noble gas has any bonds at all, it can be expanded</td>
<td></td>
</tr>
<tr>
<td>If it is not expanded valence, follow the System for Bonding by the Octet Rule (previous lesson)</td>
<td></td>
</tr>
</tbody>
</table>

**If it is expanded valence, draw the skeleton structure:**

- attach the bonding atoms to the central atom with their usual number of bonds:
- H, F, Cl, Br and I will be single bonded
- O and S will be double bonded

| 3. **Calculate the # of valence electrons we HAVE:**                   | We have:    |
| • total up the number of valence electrons each atom usually has before bonding |               |
| • add or subtract e- to account for the charge on ions                |               |
| 4. **Calculate the # of electrons that are in bonds:**                | e- in bonds: |
| • for each bond you have drawn in the skeleton structure, count 2 electrons |               |
| 5. **Calculate the # of valence electrons as lone pairs**             | # of valence electrons as lone pairs: |
| • subtract the # of electrons that are in bonds (step 4) from the # of electrons that we have (step 3) |               |
| • divide the # of electrons by 2 to find the number of lone pairs (LP) |               |
| 6. **Complete the Lewis structure:**                                  | Lewis structure: |
| • add electrons in lone pairs (step 5) around the BONDED atoms to complete stable octets for all bonded atoms (except H) |               |
| • place “left over” lone pairs on the central atom                    |               |
| • for ions, draw brackets around the finished structure and include the charge on the ion |               |
| • draw resonance structures, if necessary                            |               |

**Homework:**
2. Draw the Lewis structures for the following molecules. Use the rules for exceptions to the octet rule (insufficient and expanded valence levels).

   a) BI₃  b) ICℓ₂⁻  c) CℓF₃O₂  d) IOF₃  e) SOF₄  f) XeOF₄
   g) CℓF₄⁺  h) ICℓ₄⁻  i) XeF₂  j) BeH₂  k) IF₂O₂⁺  l) SF₄
<table>
<thead>
<tr>
<th>Total # of e- groups (BP + LP) in valence level of central atom (basic shape)</th>
<th># of bonded e-pairs (BP) around central atom (X)</th>
<th># of unbonded e-pairs (lone pairs, LP) around central atom (E)</th>
<th>General Formula (VSEPR notation) AXₙEₘ</th>
<th>Lewis Structure (use A as the central atom and X as the bonded atoms)</th>
<th>Diagram and Name of Molecular Shape</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>BeCl₂, CO₂, HCN</td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>BCl₃, CH₂O, BF₃</td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>O₃, NO₂⁻</td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>CC₁₄, NH₄⁺, ClO₄⁻¹</td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>NH₃, PF₃, ClO₂F</td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>H₂O, OF₂, I₃⁺</td>
</tr>
</tbody>
</table>

* when determining molecular shape, a double or triple bond is a single electron group. Treat multiple bonds as a single bonded pair (BP)
<table>
<thead>
<tr>
<th>Total # of e-groups (BP + LP) in valence level of central atom (basic shape)</th>
<th># of bonded e-pairs (BP) around central atom (X)</th>
<th># of unbonded e-pairs (lone pairs, LP) around central atom (E)</th>
<th>General Formula</th>
<th>Lewis Structure</th>
<th>Diagram and Name of Molecular Shape</th>
<th>Example</th>
</tr>
</thead>
</table>
| 5 | | | | | | **PCl**₅
CIF₃O₂ SOF₄ |
| 5 | | | | | | **SF₄**
IF₂O₂₁⁻ BrF₄¹⁺ |
| 5 | | | | | | **BrF₃**
ICl₃ |
| 5 | | | | | | **XeF₂**
I₃¹⁻ IC³⁺ |
| 6 | | | | | | **SF₆**
IOF₅ CIF₆¹⁺ |
| 6 | | | | | | **BrF₅**
XeOF₄ IOF₄¹⁻ |
| 6 | | | | | | **XeF₄**
BrF₄¹⁻ ICl₄¹⁻ |
To predict the shape of a molecule:
1. Draw the Lewis structure (stick diagram) for the molecule.
2. Count the number of bond pairs (BP) and lone pairs (LP) around the central atom.
3. Decide on the total number of electron groups (treat multiple bonds as single electron groups).
4. Consider the locations of lone pairs and any distortions from "regular" shapes.
5. Name the shape based on the arrangement of the bonding atoms as outlined below:

   a) If the total number of electron groups (bond pairs + lone pairs) is **TWO**:
      • two bond pairs and no lone pairs (_______): the molecule is linear
   
   b) If the total number of electron groups (bond pairs + lone pairs) is **THREE**:
      • three bond pairs and no lone pairs (_______): the molecule is trigonal planar
      • two bond pairs and one lone pair (_______): the molecule is bent or V-shaped

   c) If the total number of electron groups (bond pairs + lone pairs) is **FOUR**:
      • four bond pairs and no lone pairs (_______): the molecule is tetrahedral
      • three bond pairs and one lone pair (_______): the molecule is trigonal pyrimidal
      • two bond pairs and two lone pairs (_______): the molecule is bent or V-shaped

   d) If the total number of electron groups (bond pairs + lone pairs) is **FIVE**:
      • five bond pairs and no lone pairs (_______): the molecule is trigonal bipyramidal
      • four bond pairs and one lone pair (_______): the molecule is a “see-saw” shape
      • three bond pairs and two lone pairs (_______): the molecule is “T-shaped”
      • two bond pairs and three lone pairs (_______): the molecule will be linear

   e) If the total number of electron groups (bond pairs + lone pairs) is **SIX**:
      • six bond pairs and no lone pairs (_______): the molecule is octahedral
      • five bond pairs and one lone pair (_______): the molecule is a square-based pyramid
      • four bond pairs and two lone pairs (_______): the molecule is square planar

**Homework:**
1. Read pages 178 to 185
2. Identify the molecular shape associated with the following VSEPR notations

<table>
<thead>
<tr>
<th>a) AX₅E₁</th>
<th>e) AX₆E₀</th>
</tr>
</thead>
<tbody>
<tr>
<td>b) AX₃E₀</td>
<td>f) AX₃E₂</td>
</tr>
<tr>
<td>c) AX₄E₂</td>
<td>g) AX₃E₁</td>
</tr>
<tr>
<td>d) AX₂E₁</td>
<td>h) AX₅E₀</td>
</tr>
<tr>
<td>e) AX₄E₁</td>
<td>i) AX₄E₀</td>
</tr>
<tr>
<td>f) AX₂E₂</td>
<td>j) AX₂E₃</td>
</tr>
</tbody>
</table>

3. On pages 185 – 186, do questions 18 to 21 (refer to the examples on pages 184 to 185)
Complete the first three columns of the chart below after the lesson on Polarity of Covalent Molecules. Complete the last column after the lesson on Intra- and Inter-molecular Attraction.

<table>
<thead>
<tr>
<th>Lewis Structure</th>
<th>General Formula (AXₙEₘ) and Name of Shape</th>
<th>Determine the Polarity of the Molecule</th>
<th>Type and Strength of Inter-Molecular Forces of Attraction (IMF) and Predicted Physical Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH₃</td>
<td>Is the molecule symmetrical in all planes? _____</td>
<td>Polarity:</td>
<td>Type(s) of IMF:</td>
</tr>
<tr>
<td></td>
<td>Does the molecule contain any polar bonds? _____</td>
<td></td>
<td>Strength of IMF: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Melting point: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Solubility in water: __________</td>
</tr>
<tr>
<td>PH₃</td>
<td>Is the molecule symmetrical in all planes? _____</td>
<td>Polarity:</td>
<td>Type(s) of IMF:</td>
</tr>
<tr>
<td></td>
<td>Does the molecule contain any polar bonds? _____</td>
<td></td>
<td>Strength of IMF: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Melting point: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Solubility in water: __________</td>
</tr>
<tr>
<td>CF₄</td>
<td>Is the molecule symmetrical in all planes? _____</td>
<td>Polarity:</td>
<td>Type(s) of IMF:</td>
</tr>
<tr>
<td></td>
<td>Does the molecule contain any polar bonds? _____</td>
<td></td>
<td>Strength of IMF: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Melting point: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Solubility in water: __________</td>
</tr>
<tr>
<td>CF₃Cl</td>
<td>Is the molecule symmetrical in all planes? _____</td>
<td>Polarity:</td>
<td>Type(s) of IMF:</td>
</tr>
<tr>
<td></td>
<td>Does the molecule contain any polar bonds? _____</td>
<td></td>
<td>Strength of IMF: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Melting point: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Solubility in water: __________</td>
</tr>
<tr>
<td>CO₂</td>
<td>Is the molecule symmetrical in all planes? _____</td>
<td>Polarity:</td>
<td>Type(s) of IMF:</td>
</tr>
<tr>
<td></td>
<td>Does the molecule contain any polar bonds? _____</td>
<td></td>
<td>Strength of IMF: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Melting point: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Solubility in water: __________</td>
</tr>
<tr>
<td>SO₂</td>
<td>Is the molecule symmetrical in all planes? _____</td>
<td>Polarity:</td>
<td>Type(s) of IMF:</td>
</tr>
<tr>
<td></td>
<td>Does the molecule contain any polar bonds? _____</td>
<td></td>
<td>Strength of IMF: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Melting point: __________</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>Solubility in water: __________</td>
</tr>
</tbody>
</table>

**Homework:**
1. Read pages 187 to 189.
2. Answer questions 23 – 26 on page 188 and 1,2,3,5,6,7 on page 189.
Intra-molecular forces of attraction refer to forces of attraction _______________ molecules.
- refers to the ___________________________ that hold atoms together (ionic and covalent bonds)
- both types of bonds are very __________, but ionic bonds are _______________ than covalent bonds

Inter-molecular forces of attraction refer to the forces of attraction _______________ molecules
- these attract molecules to one another, and are __________________________ as chemical bonds
- the strength of attraction increases as the _______________ of the molecules increases
- can be used to predict the __________________________ of substances eg. __________________
- inter-molecular attractions between _____________ molecules are also known as _______________

There are three main types of inter-molecular attraction, based on the _______________ of the molecules:

1. **Ion to Ion inter-molecular attraction**
- occurs between ionic compounds, which are _______________ polar
  - positively charged ions (    ) are attracted to negatively charged  ions (    ), so the attraction is __________________________
  - the greater the charges on the ions, the ____________ the melting and boiling points
  - eg. NaCl (__________) melts at 1074 K
  - MgO (__________) melts at 3073 K
- strong inter-molecular attraction forms ________, ______________________ structures

2. **Dipole-dipole inter-molecular attraction**
   a) “Regular” dipole-dipole
   - occurs between ________________ molecules
   - permanent partial negative (     ) and permanent partial positive (     ) dipoles are attracted to each other
   - as polarity increases, the strength of the dipole-dipole interaction ________________, so melting point and boiling point ________
   - small polar molecules are often ___________ at SATP
   - large polar molecules may form ________________
   b) **Hydrogen Bonding** (a special case of dipole-dipole attraction)
   - occurs in polar molecules that contain hydrogen bonded to ____________, ____________ or ____________ (    )
   - the hydrogen atom is ____________ and carries a significant partial positive charge (    ) which is strongly attracted to the partial negative charge (    ) on other molecules
   - on average, hydrogen bonds are _____ times stronger than regular dipole-dipole attractions
H-bonding has a significant impact on physical properties such as melting and boiling points.

Tellurium, selenium, sulfur and oxygen are all in Group VI in the periodic table. The trend in the melting and boiling points for the compounds H₂Te, H₂Se and H₂S can be used to predict those of water:

The actual melting and boiling points of water are much ________ than would be predicted by periodic trends, because of _______________.

The attraction _________ water molecules is __________________, requiring more energy and higher temperatures to overcome them.

3. **Dispersion Forces (a.k.a. London Dispersion Forces)**
   - occur in ______ substances due to the random movement of _____________ in atoms
   - insignificant compared to the strength of _________ inter-molecular attraction, so _____________ in ionic substances
   - this is the only type of inter-molecular attraction between ______________ covalent compounds
   - charges are ______________ and very ________, so melting and boiling points are ______________
   - small non-polar molecules are __________ at SATP, large molecules are ___________

**Explanation:**

Random electron movement in atoms means that electrons may be __________________ clustered in one region of an atom, creating a temporary very ___________ charge. This _________ electrons in neighbouring atoms, ______________ a slight positive charge. These ___________, ____________ positive and negative charges can attract molecules weakly together.

Dispersion forces are stronger between _________ molecules since they contain more __________ to create temporary dipoles. Dispersion forces are also stronger between ____________________ or __________________ molecules because the shape of these molecules allow them to move ____________________.
# Unit 1 Summary: Structure and Properties of Substances (Chapters 3 and 4)

<table>
<thead>
<tr>
<th>Topic</th>
<th>Key Points</th>
</tr>
</thead>
</table>
| **Atomic Models (historically)** | • Contributions of Rutherford, Bohr, Planck, de Broglie, Einstein, Schrodinger  
• “the basics”: atomic number, mass number, charge on ions, isotopes, eg. Ag-107 |
| **Quantum Mechanical Model** | • aufbau principle, Hund’s rule, Pauli exclusion principle  
• quantum mechanical model of the atom summary sheet |
| **Electron configurations, orbital diagrams and quantum numbers** | • read, write and interpret e- configurations  
• condensed e- configurations (use the Noble gas from the *previous* period, eg. F⁻¹ [He]²s²²p⁶)  
• predicted vs. actual e- configurations (sp¹, sp², sp³ etc. hybrids, filling and half-filling the d-block orbitals, eg. Ag, Cr, Cu etc) |
| **Trends on the P.T.** | • $Z_{eff}$ and shielding effect  
• EN, IE, EA, atomic radius, ionic radius, metallic character, reactivity of metals and non-metals |
| **Bonding (Intra-molecular Forces of Attraction)** | • type of bonding depends on EN and IE of the atoms involved  
• be able to explain electron arrangement in ionic, pure and polar covalent and metallic bonding  
• classify *bonds* using $\Delta$EN and % ionic character |
| **Lewis Structures and AXₙEₘ** | • VSEPR theory to predict shapes of molecules  
• VSEPR notation ($AX_nE_m$)  
• bond angles of basic shapes of molecules  
• symmetry of molecules |
| **Polarity of Molecules** | • depends on: shape and symmetry of the molecule and the polarity of bonds |
| **Inter-molecular Forces of Attraction** | • distinguish between intra-molecular forces (bonds) and inter-molecular forces  
• strength of inter-molecular forces (ionic > hydrogen bonding > dipole-dipole > dispersion forces) |
| **Properties of substances** | • determined by the strength of intra and inter-molecular forces  
• types of crystalline solids and their properties  
• be able to *explain* the properties of crystalline solids in terms of the intra and inter-molecular forces (eg. melting point, boiling point, solubility, odour, hardness) |