1. Molecule  | bond length  | X-H (Å)  
NH₃        | 1.02        
PtH₂       | 1.42        
SbH₃       | 1.71        

As we move down the group, the X-H bond length increases. 
AsH₃ will therefore have a As-H bond length between 
P-H and Sb-H. In the absence of any other information, 
let's assume that the As-H bond length is the average 
of P-H and Sb-H: 

\[ \text{As-H} = 1.57 \, \text{Å} \]  
(measured value is \(1.52 \, \text{Å}\))

The longest bond, Sb-H, is the weakest.

2. We see from Table 3.3 that F₂ has an unexpectedly weak bond 
compared to the other halides (Cl₂, Br₂, I₂). Because we 
also see from Table 3.3 that higher bond lengths are correlated 
with lower bond energies, this implies that the longer than expected 
F₂ bond length is because it is a weaker than expected bond. We can speculate 
that this is because of the F atoms' unique EN.
3. Fraction ionic character = \( f = 0.208 \frac{D}{D^*} \) \\
\( D^* = \text{dipole moment (given)} \) \\
\( D = \text{bond length (given)} \) \\

So for each in units of D and r in units of Å, we will get a unitless fractional ionic character.

Fill out the table:

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Bond length (Å)</th>
<th>Dipole moment (D)</th>
<th>f</th>
<th>%</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO</td>
<td>1.573</td>
<td>1.234</td>
<td>0.16</td>
<td>16</td>
</tr>
<tr>
<td>KI</td>
<td>3.051</td>
<td>10.82</td>
<td>0.74</td>
<td>74</td>
</tr>
<tr>
<td>TICl</td>
<td>2.408</td>
<td>4.543</td>
<td>0.38</td>
<td>38</td>
</tr>
<tr>
<td>InCl</td>
<td>2.404</td>
<td>3.79</td>
<td>0.33</td>
<td>33</td>
</tr>
</tbody>
</table>

For these 4 molecules, the molecule ClO has the greatest EN difference between the two ions (KI) has the highest fractional ionic character, while the molecule ClO has the lowest EN difference (ClO) has the lowest fractional ionic character. This makes sense.
4. a) \( \text{ICl}_4^- \)

\[
\begin{align*}
4 \times \text{Cl} &= 28 \text{e}^- \\
\text{I} &= 7 \text{e}^- \\
\text{E} &= 1 \text{e}^- \\
\hline
36 \text{e}^- 
\end{align*}
\]

First draw Lewis structure:

So I has 4 bonded atoms and 2 lone pairs.

\[\text{SN} = 6\]

Now decide geometry: Molecule is octahedral \((\text{SN}=6)\)

Where do the lone pairs go?

All interactions are 90° apart, so unlike the trigonal bipyramidal structure, there is no obvious place the lone pairs must go. However, once we put 1 lone pair on, we need to minimize its interaction with the other lone pair, so the two should be 180° apart.

\[\text{Cl} \text{ Cl} \text{ Cl} \text{ Cl} \quad \text{or} \quad \text{Cl} \text{ Cl} \text{ Cl} \]

would both be acceptable.
b) \( \text{OF}_2 \):
\[
\begin{align*}
\text{SN} &= 4 \\
\text{tetrahedral, w/a "bent" Structure}
\end{align*}
\]

\( c) \text{BrO}_3^- \)
\[
\begin{align*}
\text{BR} &= 7e^- \\
\text{O} &= 18e^- \\
\text{O} &= 1e^- \\
\text{O} &= 26e^-
\end{align*}
\]

\[
\begin{align*}
\text{SN} &= 4, \text{tetrahedral w/a} \\
\text{"pyramidal" Structure}
\end{align*}
\]

\( d) \text{CS}_2 \)
\[
\begin{align*}
\text{SN} &= 2, \text{linear}
\end{align*}
\]

\( e) \text{CHO}_3^- \)
\[
\begin{align*}
\text{SN} &= 3, \text{trigonal planar}
\end{align*}
\]
S. 

\[
\begin{align*}
\text{N} & \quad \text{O} \\
\text{N} & \quad \text{O} \\
\text{N} & \quad \text{O}
\end{align*}
\]

\[
\begin{align*}
\text{N} \quad \text{O} \quad \text{O} & \Rightarrow \quad \text{N} \quad \text{N} \quad \text{O} \\
\text{N} \quad \text{N} \quad \text{O} & \Rightarrow \quad \text{N} \quad \text{N} \quad \text{O}
\end{align*}
\]

Both resonance structures are linear.

Although these are both technically correct Lewis structures, they have very different arrangement of formal charges. We should always try to put formal charge on the least electronegative atom, so the 3rd structure is "better" in that it makes more chemical sense. If dipole points from negative to positive, then

\[
\begin{align*}
\text{N} & \quad \text{N} \quad \text{O} \\
\text{N} & \quad \text{N} \quad \text{O}
\end{align*}
\]

6. The shorter the bond length, the closer two nuclei are held together. To overcome the electrostatic repulsion between the two nuclei, this means that the bonding, attractive interaction caused by the shared electrons must be greater. Therefore, in general, shorter bonds are associated with higher bond dissociation energies (i.e. stronger bonds).
7. Octahedral w/ 0 lone pairs:

Each Y has interacts w/ 4 other Y's 90° away and
1 Y 180° away. Therefore all Y's are structurally identical.
Therefore, we can add our 1st lone pair anywhere.

1 lone pair:

Now the molecule is structurally asymmetric. We need to add
the 2nd lone pair as far away from the 1st as possible

2 lone pairs:

All Y's are now structurally identical again, doesn't matter
where we add the 3rd lone pair.

3 lone pairs:
8. \( \text{XeF}_2 \):

\[
\begin{align*}
\text{Xe} &= 8e^- \\
2\text{e}^- &\quad \text{XeF} = 14e^- \\
&\quad \text{Xe}^+ \\
&\quad \text{F}^-
\end{align*}
\]

- \( \text{F} - \text{Xe}^+ - \text{F} \):
  - \( \text{SN} = 5 \), 3 lone pairs to deal with

- \( \text{F} - \text{Xe} - \text{F} \):
  - Triangular bipyramidal, all lone pairs in the same plane

\( \text{XeF}_4 \):

\[
\begin{align*}
\text{Xe} &= 8e^- \\
4\text{e}^- &\quad \text{XeF} = 18e^- \\
&\quad \text{Xe}^+ \\
&\quad \text{F}^-
\end{align*}
\]

- \( \text{F} - \text{Xe}^+ - \text{F} \):
  - \( \text{SN} = 6 \), 2 lone pairs

- \( \text{F} - \text{Xe} - \text{F} \):
  - Square planar
\[ \begin{align*}
\text{XeO}_3 : & \quad \text{Xe} = 8e^- \\
& \quad 3\text{Xe} + 3\text{O} = 18e^- \\
& \quad \text{SN} = 4, \text{ 1 lone pair}
\end{align*} \]

\[ \begin{align*}
\text{XeO}_4 : & \quad \text{Xe} = 8e^- \\
& \quad 4\text{Xe} + 4\text{O} = 32e^- \\
& \quad \text{SN} = 4, \text{  no lone pairs}
\end{align*} \]

\[ \begin{align*}
\text{H}_4\text{XeO}_6 : & \quad \text{Xe} = 8e^- \\
& \quad 6\text{Xe} + 6\text{O} = 36e^- \\
& \quad 4\text{H} + 4e^- = 48e^- \\
& \quad \text{SN} = 6, \text{ 0 lone pairs}
\end{align*} \]
$XeOF_4$

$Xe = 8e^-$
$O = 6e^-$
$4F = 25e^-$

$42e^-$

$SN = 4$, lone pair