20. a. $Rb < K < Na$ b. $Ga < B < O$ c. $Br < Cl < F$ d. $S < O < F$

26. Ionic character is proportional to the difference in electronegativity values between the two elements forming the bond. Using the trend in electronegativity, the order will be:

Br–Br	<	N–O	<	C–F	<	Ca–O	<	K–F
smallest ionic character								most ionic character

Note that Br–Br, N–O and C–F bonds are all covalent bonds since the elements are all nonmetals. The Ca–O and K–F bonds are ionic as is generally the case when a metal forms a bond with a nonmetal.

36. a. $V > V^{2+} > V^{3+} > V^{5+}$ b. $Cs^+ > Rb^+ > K^+ > Na^+$ c. $Te^{2-} > I^- > Cs^+ > Ba^{2+}$

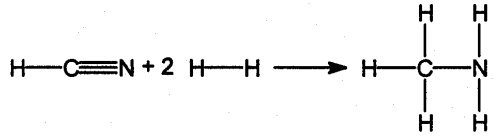
d. $P^{3-} > P^{2-} > P^- > P$ e. $Te^{2-} > Se^{2-} > S^{2-} > O^{2-}$

32. a. V^{5+} b. none c. Sr^{2+} d. P^{3-}

42.

$Mg(s) \rightarrow Mg(g)$	$\Delta H = 150.\ kJ$	(sublimation)
$Mg(g) \rightarrow Mg^+(g) + e^-$	$\Delta H = 735\ kJ$	(IE_1)
$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$	$\Delta H = 1445\ kJ$	(IE_2)
$F_2(g) \rightarrow 2\ F(g)$	$\Delta H = 154\ kJ$	(BE)
$2\ F(g) + 2\ e^- \rightarrow 2\ F^-(g)$	$\Delta H = 2(-328)\ kJ$	(EA)
$Mg^{2+}(g) + 2\ F^-(g) \rightarrow MgF_2(s)$	$\Delta H = -3916\ kJ$	(LE)
$Mg(s) + F_2(g) \rightarrow MgF_2(s)$	$\Delta H_f^\circ = -2088\ kJ/mol$	

48. Sometimes some of the bonds remain the same between reactants and products. To save time, only break and form bonds that are involved in the reaction.

a. 

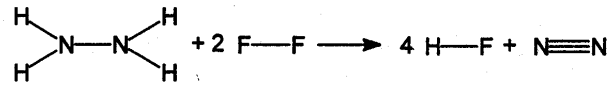
Bonds broken:

1 C ≡ N (891 kJ/mol)
2 H - H (432 kJ/mol)

Bonds formed:

1 C - N (305 kJ/mol)
2 C - H (413 kJ/mol)
2 N - H (391 kJ/mol)

$\Delta H = 891\ kJ + 2(432\ kJ) - [305\ kJ + 2(413\ kJ) + 2(391\ kJ)] = -158\ kJ$

b. 

Bonds broken:

1 N - N (160. kJ/mol)
4 N - H (391 kJ/mol)
2 F - F (154 kJ/mol)

Bonds formed:

4 H - F (565 kJ/mol)
1 N ≡ N (941 kJ/mol)

$\Delta H = 160.\ kJ + 4(391\ kJ) + 2(154\ kJ) - [4(565\ kJ) + 941\ kJ] = -1169\ kJ$

60. $NH_3(g) \rightarrow N(g) + 3\ H(g)$; $\Delta H° = 3\ D_{NH} = 472.7 + 3(216.0) - (-46.1) = 1166.8$ kJ

$$D_{NH} = \frac{1166.8\ \text{kJ}}{3\ \text{mol NH bonds}} = 388.93\ \text{kJ/mol} \approx 389\ \text{kJ/mol}$$

$D_{calc} = 389$ kJ/mol as compared to 391 kJ/mol in the table. There is good agreement.

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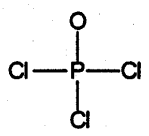
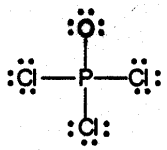
64. a. $POCl_3$ has $5 + 6 + 3(7) = 32$ valence electrons.



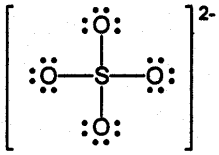
Skeletal structure



Lewis structure

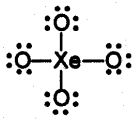
This structure uses all 32 e^- while satisfying the octet rule for all atoms. This is a valid Lewis structure.

SO_4^{2-} has $6 + 4(6) + 2 = 32$ valence electrons.

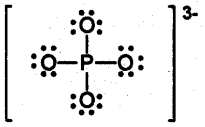


Note: A negatively charged ion will have additional electrons to those that come from the valence shells of the atoms.

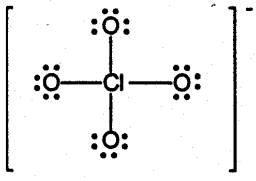
XeO_4, $8 + 4(6) = 32\ e^-$



PO_4^{3-}, $5 + 4(6) + 3 = 32\ e^-$



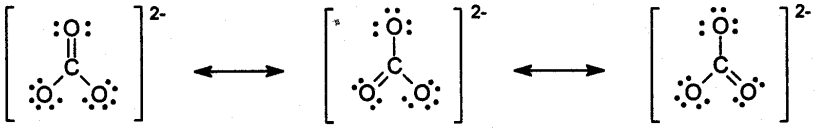
ClO_4^- has $7 + 4(6) + 1 = 32$ valence electrons.



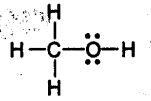
74. The Lewis structures for the various species are below:

CO (10 e⁻): :C≡O: Triple bond between C and O

CO_2 (16 e⁻): O=C=O Double bond between C and O

CO_3^{2-} (24 e⁻): 

Average of 1 1/3 bond between C and O

CH_3OH (14 e⁻):  Single bond between C and O

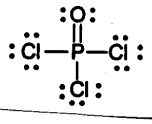
As the number of bonds increase between two atoms, bond length decreases and bond strength increases. With this in mind, then:

longest → shortest C – O bond: $CH_3OH > CO_3^{2-} > CO_2 > CO$

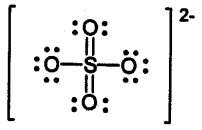
weakest → strongest C – O bond: $\mathbf{CH_3OH < CO_3^{2-} < CO_2 < CO}$

For SO_4^{2-}, ClO_4^-, PO_4^{3-} and ClO_3^-, only one of the possible resonance structures is drawn.

a. Must have five bonds to P to minimize formal charge of P. The best choice is to form a double bond to O since this will give O a formal charge of zero and single bonds to Cl for the same reason.

 P, FC = 0

b. Must form six bonds to S to minimize formal charge of S.

 S, FC = 0

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80. See Exercises 8.64 and 8.68 for the Lewis structures.

8.64 a. All are tetrahedral; 109.5°

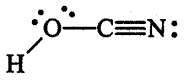
b. All are trigonal pyramid; < 109.5°

c. All are V-shaped; < 109.5°

8.68 O_3 and SO_2 are V-shaped (or bent) with a bond angle ≈ 120°. SO_3 is trigonal planar with 120° bond angles.

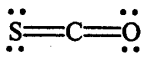
88. All have polar bonds; in SiF_4 the individual bond dipoles cancel when summed together, and in PCl_3 and SCl_2 the individual bond dipoles do not cancel. Therefore, SiF_4 has no dipole moment (is nonpolar) and PCl_3 and SCl_2 have dipole moments (are polar). For PCl_3, the negative end of the dipole moment is between the more electronegative chlorine atoms and the positive end is around P. For SCl_2, the negative end is between the more electronegative Cl atoms and the positive end of the dipole moment is around S.

92. a.



Polar; The bond dipoles do not cancel.

b.



Polar; The C – O bond is a more polar bond than the C – S bond. So the two bond dipoles do not cancel each other.